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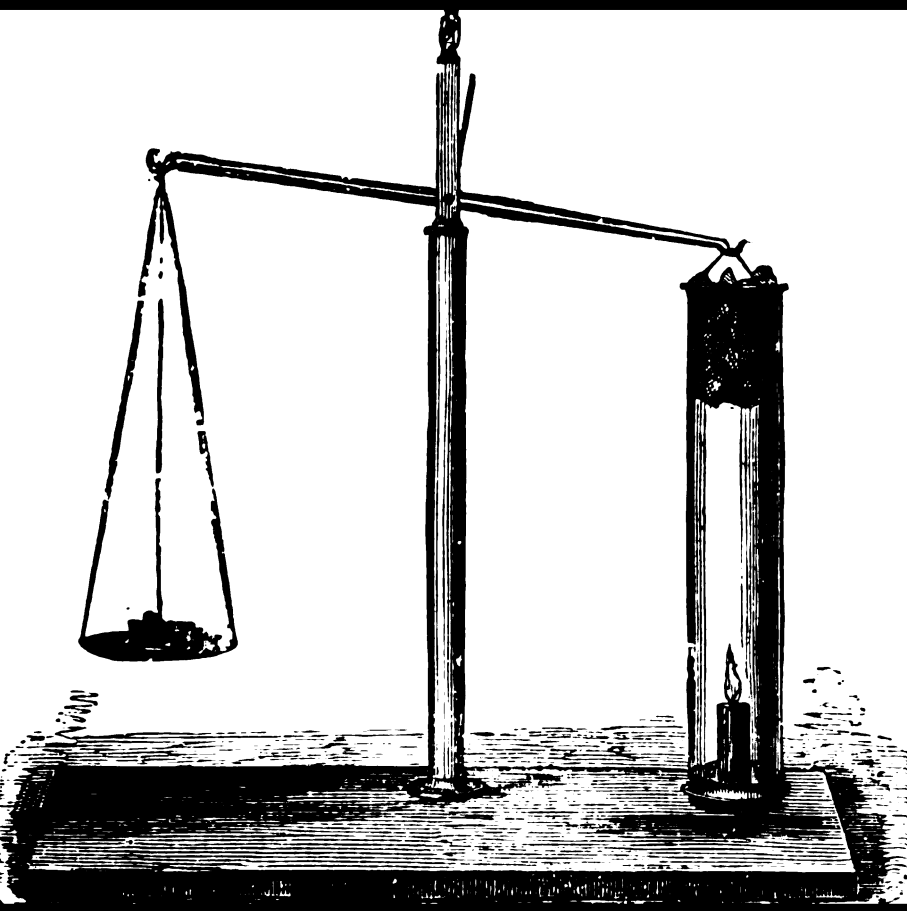

Chemistry of common objects adapted to the alternative ...

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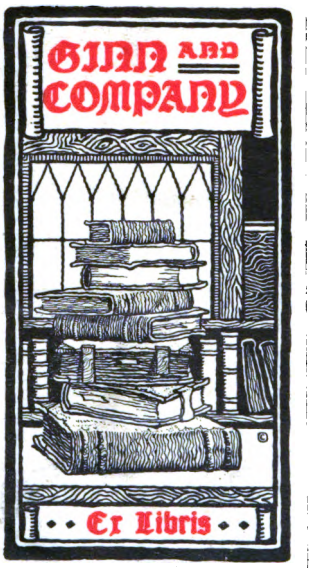

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THIRTY ELEMENTARY LESSONS

ON THE

CHEMISTRY

OF

COMMON OBJECTS.

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"PROGRESSIVE QUESTIONS IN CHEMISTRY"—6d.

This little work contains all the questions set at the Examinations of the Department of Science and Art in Inorganic Chemistry for the past twenty years. The questions are progressively arranged, and they will be found of great assistance in directing the studies of those preparing for examinations.

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CHEMISTRY

OF

COMMON OBJECTS

ADAPTED TO THE

Alternative Elementary Stage of the Syllabus of the Department of Science and Art.

BY

JOHN J. PILLEY.

London:

GEORGE GILL & SONS,

13, WARWICK LANE, E.C.

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PREFACE.

THESE short outlines of the chemistry of common objects formed originally a course of lectures to working men, given under the auspices of the Trustees of the St. Luke's Parochial Charities, London. They were mainly modelled upon the lines of the excellent syllabus published by the Department of Science and Art. Since they were first delivered they have been repeated, in a modified form, to large classes of school teachers; and it is the demand for a simple guide to the chemistry of familiar objects for such classes, which has led to their publication in the present form. In preparing them for press I have found it advisable to make very extensive use of the particular form of phraseology adopted in the syllabus published for the *Alternative Elementary stage* by the Department of Science and Art. All the subjects mentioned in that course are fully treated of, but in order to increase the value of the work for practical purposes, I have found it advisable to give three introductory lessons on matter and energy, chemical action, and elements and compounds; for the same reason I have, in some other respects, gone beyond the actual verbal requirements of the syllabus. This, I believe, makes the work more comprehensive in scope and better adapted as an introduction to the study of *Agriculture*, *Hygiene*, *Physics*, *Physiography*, *Physiology*, *Geology*, and the other branches of *Natural Science*.

The most one can aim at in preparing such an elementary work is simplicity of method and exactitude of matter; these objects I

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have kept well in view, and have endeavoured to put the facts into such a form that they will prove both a simple and reliable guide.

The omission of all chemical formulæ ought to make the work both readable and intelligible for the general student; and the following special features should commend it to, at least, the serious attention of teachers :—

1. It contains an outline of the chemistry of common objects, embracing completely the whole of the *Alternative Course of the Department of Science and Art, in Thirty Lessons.*
2. The "apparatus required" and "experiments to be performed" in the lessons are given in tabular form before each chapter; these should prove time-saving both to the teacher and student.
3. Each chapter is followed by a "*Summary*" which contains all the important facts which should be on the blackboard of the teacher and in the notebook and mind of the student at the close of the lesson. These should prove valuable alike to instructor and instructed.
4. The sets of "*Questions*" placed at the end of each chapter are intended to test the class work.
5. The *experimental method* adopted should make the work of special educational value to young students, since by it they are taught what to do, how to observe, and are instructed in inductive and deductive methods of reasoning.

Elementary school teachers will find the subjects treated of well adapted for a course of *Object Lessons.*

I have, in conclusion, to acknowledge my indebtedness to the authorities who are responsible for the syllabus published by the Science and Art Department, which syllabus has been my chief guide. My sincere thanks are also due to my colleague, Mr W. N. Edwards, F.C.S., for his valuable assistance in reading the proofs.

J. J. P.

Dulwich, September 1892.

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Preparation for First Lesson.

APPARATUS AND MATERIALS REQUIRED.

Ice, platinum wire, pieces of soft iron and steel, rods of ebonite, sealing wax, amber, and sulphur, piece of flannel, rusted iron, sulphuric acid, matches, iron filings, magnesium wire, iron poker, Bunsen burner or spirit lamp, specimens of granite, chalk, clay, sand, marble, slate, coal, common ores, graphite, sulphur, turpentine, glycerine, and oil.

EXPERIMENTS TO BE PERFORMED.

Exhibit to the class and pass round for examination specimens of rocks, minerals, ores, and various liquids. Call attention to marked differences.

Show that ice melts on the application of heat, and that steam condenses to water.

Show that platinum wire when heated gives out light.

Magnetise a piece of iron and show that it will attract iron filings, and that it acts upon a magnetic needle.

Electrify pieces of sealing wax, amber, ebonite, etc., and show how they possess new properties.

Ignite a piece of phosphorus by friction, and repeat the experiment with the top of a friction match.

Burn a piece of magnesium ribbon and some magnesium dust, and show that the bodies are entirely changed.

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THE CHEMISTRY
OF
COMMON OBJECTS.

CHAPTER I.

MATTER AND ENERGY.

Chemistry—Physical Science—Matter—Physical and Chemical Changes—Common Examples of Chemical and Physical Changes—Relationship between Physics and Chemistry.

CHEMISTRY is the name given to one branch of natural science, the special object of which is the discovery of the composition of all things. We are all familiar with the fact that we are surrounded by an infinite variety of different kinds of substances which possess very distinctive properties by which we can identify them. The **rocks** composing the solid earth, such as *sandstone*, *chalk*, *slate*, *clay*, *granite*, *limestone*, and *marble*; the **minerals**, like *coal*, *graphite*, *sulphur*; and the **ores** from which metals are extracted, are all materials which differ very much indeed from each other. Again, the common **liquids**—*water*, *oil*, *turpentine*, *benzoline*, and *glycerine*—all possess very marked properties by which we are able with ease to recognise any one of them. *Coal gas* and *air* are both **gases**, yet they differ so much from one another, in certain

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respects, that we have not the least trouble about identifying each. In the great kingdom of living things the various kinds of bodies present in and obtainable from both plants and animals are practically infinite. Since the chief object of chemistry is to discover what all the many different kinds of substances are made of, it follows that a knowledge of this branch of science must prove a very valuable assistance to the student of any other section of natural science. The student of **Botany,** for instance, will better understand the nature of the components of plants, how air assists the plant to grow, and how the substances produced in vegetables are fitted for foods for animals. One who begins the study of **Physiology,** with a knowledge of chemistry, will find little difficulty in understanding how it is that food stuffs are changed and made ready to enter the blood, how the blood takes up oxygen from the air we breathe, and what becomes of that oxygen. These and many kindred problems can only be clearly understood by those who have at least an elementary knowledge of chemistry. So also the study of this branch of science is more or less closely related to all other branches of natural science. Since chemistry aims at discovering the composition of *all* things, it must have something to tell us about water and air, and the composition of everything which may be got out of the earth and obtained from plants and animals.

NATURAL SCIENCE.—The word science means knowledge, and the exact knowledge of anything is therefore a science; but there are many kinds of knowledge, and in

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these lessons we hope to be able to learn a little of one branch of *natural* science. Remember that natural science may be described as the study of the objects of nature and the changes which they undergo. The objects of nature are of course all those things which exist in the visible world around us; and the changes which they undergo give rise to the phenomena of light, heat, the winds, etc. It aims, therefore, at teaching us something of the earth on which we live, the water which covers so large a portion of its surface, the air by which it is surrounded, the many plants and animals found upon its surface, and of the far-distant heavenly bodies.

Since the study of natural objects is so very wide, and as our knowledge of things and causes is continually increasing, it has been found convenient to divide the whole. Thus one section dealing with living things is termed **Biology**; whilst that which treats of the past history and probable future of the earth is **Geology**; the study of the heavenly bodies is **Astronomy**, and the knowledge of the laws regulating the motion of things **Mechanics**.

MATTER.—All that which exists and is capable of, in some way, affecting our senses is termed matter. It may be visible like solid earth, or invisible as air and coal gas,—near like this book, or far away as the sun and moon. Anything which we can see or feel or of which we can learn something by handling and examination is matter, substance, or stuff. Remember that all things, the air we breathe, the food we eat, the water we drink, the clothes we wear, the tools we employ, the coal and gas which we burn, and

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the many substances which men get out of the earth and all the components of plants and animals, are forms of matter. And all the knowledge we have of these things has been gained by examination and handling—that is, by observation and experiment; and the **record of this knowledge is natural science.**

When we consider the above facts we see that this stuff or matter exists in three well-known conditions or states: **solid,** such as *glass, sand, coal, paper, iron,* and *sulphur*; **liquid,** like *water, turpentine,* and *glycerine*; **gaseous,** for example, *air* and *coal gas.* The various forms of matter can, however, all be grouped into two classes, namely, solids and fluids.

A Solid is a body which offers resistance both to change of form and change of bulk.

A Fluid offers no resistance to change of shape; *gases, vapours,* and *liquids* are all fluids.

All kinds of matter are made up of minute fragments which are more or less closely bound together, and to which they may be reduced by physical means; these component fragments are termed **particles.** The force which holds these particles together is termed **cohesion,** and the several conditions or states of matter depend upon the relationship which exists between the ultimate particles. In the solid state the force of cohesion is greatest. In liquids it is less, for in this case the particles freely glide round each other. In the gaseous condition of matter the particles are self-repellent; for in such bodies the force of cohesion is overcome by another force called **repulsion,** and the particles are driven apart.

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ENERGY.—Matter is constantly undergoing changes: for example, iron rusts when exposed to air; coal burns when heated; a bell emits a sound when struck; a raindrop falls to the earth; water dries up and disappears. Matter is in an everlasting state of unrest, and thus its changes give rise to all the phenomena which occur in the world around us.

Now this continual change of condition and position is said to be due to a something which belongs to matter, which students of natural science call energy. It follows from the above that, if all kinds of substances or stuff which we know of are forms of matter, and the manifold changes which it undergoes are due to energy, *matter and energy are the causes of all the phenomena of nature.* In the study of natural science we endeavour to learn something of these things.

PHYSICAL CHANGES OF MATTER.—We are all familiar with many examples of changes which matter undergoes, but students of natural science find it convenient to group these changes under two heads, the first of which is given above. When water is exposed to warm air it undergoes a change, for it disappears. When cold iron is thrust into the fire it becomes hot, and may after a time become so greatly changed that it gives out both light and heat. A piece of ice placed in the teakettle is changed by heat into water, then into steam, and by cooling, the steam may be reconverted into water, and the water by further loss of heat once again altered into ice. All these changes may be brought about without any

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increase or decrease of weight, so that the total quantity of matter remains unaltered, although the condition of the substance is changed. Whatever, then, the nature of the change, since there is neither loss nor gain in weight, we may speak of it as non-material.

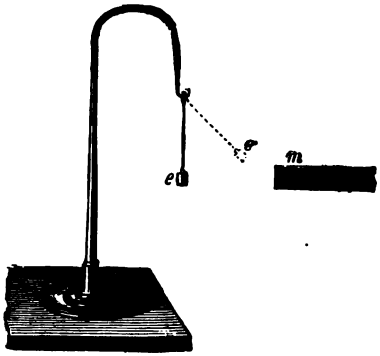

Fig. 1.—Iron attracted by a magnet.

Experiment 1.—Place a piece of *platinum* wire, about two inches in length, in the non-luminous flame of a Bunsen burner; after a short time withdraw the wire, and allow it to cool.

Observe that the metallic platinum becomes red hot, glows and gives out a bright light, and that when cool it assumes its original bright white condition, and proves in every respect to be quite unaltered.

Fig. 2.—Magnetised iron attracting iron filings.

We learn from this that the piece of matter has undergone a change which was only *temporary*, and in no sense material; for its characteristic properties are not modified, nor is there any alteration in weight, and this change was produced by the action of **heat.**

When a piece of steel is stroked with a strong magnet it will afterwards pick up light pieces of iron, and thus we learn that it has been made by this means to acquire new properties by which it may be distinguished from ordinary unmagnetised iron. The force which has given rise to this changed condition of the iron is termed **magnetism.**

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Experiment 2.—Bring a piece of iron or a knife blade near to some iron filings, and observe that the iron filings are not attracted; now draw a magnet along the piece of iron or knife blade some twenty or thirty times in the same direction, and then bring it again near to some fine pieces of iron, or a small piece of soft iron suspended, as in Fig. 1.

Observe that the filings are now attracted by the magnetised piece of iron, and that the suspended iron moves towards the magnet.

We learn from this that the metal iron has certainly undergone a change, which we may prove is not material, for the weight and appearance remain quite unaltered, and all the ordinary properties by which we recognise iron are possessed by the magnetised iron; and further, after a time, it loses its magnetism; or, if strongly heated, it ceases at once to be a magnet. Thus, as with the red-hot metal in our last experiment, here again we have an example of **a real change of matter**, for through magnetism it is endowed with new powers, just as in the previous case, through heat, it was made to exhibit new properties.

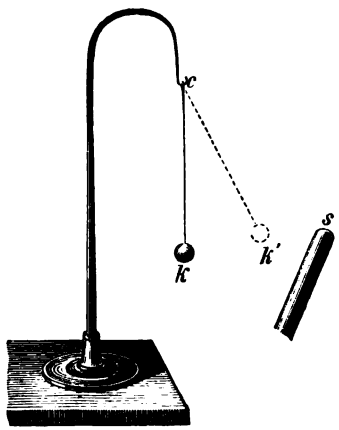

Fig. 3.—Pith-ball attracted by an electrified rod.

Many bodies when rubbed acquire, temporarily, new properties; thus dry warm sulphur, brown paper, indiarubber, amber, sealing-wax, etc., when rubbed will pick up light substances, and exhibit other interesting phenomena.

Experiment 3.—Bring a dry warm stick of *sealing-wax* and *piece of flannel* near to some fragments of paper, straw, or sawdust. Now briskly rub the rod with the flannel, and then bring it once again near to the light materials.

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Observe that in the first case the light fragments are not acted upon by the rod, but that after rubbing, the sealing-wax attracts and then repels the pieces of straw or sawdust.

We learn from this that the substance has become possessed of some new properties; it has most certainly undergone some change as a consequence of the rubbing, for it now exhibits quite new phenomena. Yet the weight of the rod remains unaltered, and it is still sealing-wax and nothing else, and, what is more, after a few minutes it loses its newly-acquired powers. The changed condition is therefore not material, nor is it in any degree permanent. The force which has given rise to this change in matter is known as **electricity.**

In the cases studied in experiments 1, 2, and 3, the bodies employed have gained certain new powers, and exhibited new phenomena, but they have not undergone any material change, and in each case those properties by which we should ordinarily identify these different kinds of matter would enable us still to recognise them at once.

The red-hot poker becomes dull as it cools, and finally is as cool and non-luminous as surrounding objects. Heat will change solid ice into liquid water, and the loss of heat causes gaseous steam to pass into liquid water; now the water produced from melting ice is exactly the same as the water obtained from condensed steam. And what is more, one pound of ice will produce exactly one pound of water, and one pound of steam will also yield exactly one pound of water; so in the change of condition or state there is no alteration of the total quantity of matter. Ice, water, and steam are all distinguished by the same general properties; they will not burn, but put out fire. It should

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be clear then to the student that the changes of ice into water, and of water into steam, are not really material or permanent changes, but merely **temporary alterations of state** or condition dependent upon temperature.

All such changes as those which we have studied above, in which matter acquires for a time new properties, and is comparatively temporarily altered, are known as **physical changes**, and the special study of the causes of such changes and the laws regulating them is termed **physics**.

CHEMICAL CHANGES OF MATTER.—Other kinds of changes than those above described are ever taking place in the world around us. These are all attended by real and permanent modifications in the properties of the matter concerned. Examples of such changes must be familiar to all. We know how metals tarnish if exposed to damp air, and every reader of this must have seen how bright metallic iron under such conditions rusts ; and some have observed how very unlike the bright hard metal is the soft brownish powder which is produced upon its surface. When milk turns sour, or the juice of the grape ferments and yields wine, and when this wine changes into vinegar, and when vegetable and animal matters decay, in each case bodies are produced with new properties, which differ widely from the substances which yield them. Sour milk is very unlike fresh milk ; wine differs both from the grape juice which yields it and from vinegar which it passes into ; the green leaf of a tree is quite unlike the substances which are produced during its decay, and the piece of meat which is allowed to rot yields bodies entirely unlike it. In all these cases we have really examples of

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material changes. Again, when substances burn they disappear, and one may be inclined at first to say that the matter has been destroyed; but we shall soon learn that **chemistry teaches us that matter cannot be really lost or destroyed**, although bodies may be made to undergo very great changes. Because a substance changes its form, even from a visible solid into invisible gases, we must not be tempted into the belief that the matter of which it is composed is destroyed; but we must at all times bear in mind the fact that matter is in an everlasting state of change, and that there is no such thing as real permanency of form or condition. The matter of which fuel, oil, and candles consist is not really lost or destroyed when these bodies are burnt, for new kinds of substances are produced, and the matter still exists, although in a new form.

Such *comparatively* permanent and material changes as those described above are due to alterations in the composition of matter, and are known as **chemical changes**, and the study of the causes of such changes is **chemistry**. We must learn now, by means of a few experiments, to distinguish a physical from a chemical change.

Experiment 4.—Hold a piece of *magnesium* wire, about two inches in length, in the Bunsen flame by means of a pair of tongs.

Observe that it becomes red hot and bursts into flame (compare experiment 1), and burns with a brilliant light, white fumes float away, and a fine white powder falls on the bench; this powder is *magnesia*.

We learn from this that the metal magnesium when heated, unlike the platinum, undergoes a real material change, and that the resulting white earthy powder is entirely unlike the bright metallic magnesium; this therefore is an example of a chemical change.

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Experiment 5.—Cut the red top off an ordinary *friction match*, and place it between two small pieces of sandpaper, and then rub the two pieces of paper together.

Observe that the match top takes fire, and in burning gives rise to strong-smelling vapours.

We learn from this that the red material from the top of the match has undergone a chemical change, because the bodies produced are altogether unlike the substance used.

Such changes as those described above, are due to a force acting in matter which is known as **Chemical Affinity.**

CHEMISTRY AND PHYSICS.—The study of those changes in the composition of matter all of which give rise to permanent alterations of the properties of bodies are included under the head of **Chemistry.** The branch of science which treats of the changes taking place in matter which give rise to sound, light, heat, electricity, magnetism, and motion, is known as **Physics.** But these branches of study are very closely related; so dependent are they one upon the other, that we cannot learn anything of either alone. As we proceed with our studies, we shall find that the phenomena which each attempts to discover the causes of are very closely related. We shall see that chemical action may be made to excite electricity, and this in its turn may give rise to light and heat and magnetism, and these again to sound and motion. Heat, electricity, magnetism, and light may each in their turn give rise to chemical action. We should bear well in mind that all changes of matter, whether we call them for convenience chemical or physical, are due to energy, and that it is this **ever-acting energy which gives rise to all the phenomena of nature.**

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SUMMARY TO CHAPTER I.

CHEMISTRY - - - The *branch of natural science* which aims at discovering the composition of all things which exist in the world around us.

NATURAL SCIENCE The exact study of the objects of nature, and the causes of the phenomena exhibited by them.

PHENOMENA OF NATURE DUE TO -
Matter, that which exists; known in the solid, liquid, and gaseous states; composed of particles held together by the force of cohesion.
Energy, that which gives rise to changes in position or condition of matter. It is the cause of light, heat, magnetism, gravitation, electricity, and motion.

TWO GROUPS OF CHANGES AND PROPERTIES OF MATTER - - -
Physical changes, in which bodies undergo *no alteration of weight*, but *temporarily* acquire distinctly new properties; may be brought about by *heat*, *light* *electricity*, and *magnetism*, etc.
Chemical changes of matter are attended by *an alteration in weight*, and the substance becomes comparatively *permanently altered* in character. Such changes are due to the action of *chemical affinity*.
The Physical properties of a body are those which may be discovered without employing chemical means and appliances; they include colour, taste, odour, lustre, fracture, specific gravity, hardness and physical state, and whether soluble or insoluble, in certain liquids.
The Chemical properties of a substance must be ascertained by strictly chemical methods; these include behaviour with certain reagents, and the compounds the body is capable of forming.

PHYSICS AND CHEMISTRY - - Closely related, but the changes in condition of matter due to heat, light, electricity, and magnetism, studied as *Physics*, whilst the study of the more permanent changes brought about through alterations in composition are included in *Chemistry*. But really only one kind of *energy* giving rise to all phenomena, whether they are spoken of as chemical or physical.

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QUESTIONS ON CHAPTER I.

1. What does chemistry teach us about?

2. What do you understand by natural science?

3. Name some of the branches of natural science, and state what each teaches us about.

4. What is matter? In how many states does matter exist? Give four examples of each condition.

5. What is the difference between a solid and a fluid? What classes of bodies are included under the head of fluids?

6. How are the particles of bodies held together, and what is the difference in this particular between solids, liquids, and gases?

7. What is energy? Name some natural phenomena which are due to energy acting in matter.

8. What are the chief characters of physical changes of matter? Give three examples of physical changes.

9. Describe fully any experiment which you have performed which demonstrates the nature of a physical change.

10. Suggest an experiment which you consider would illustrate a physical change.

11. Show how a physical change may be produced in matter by heat.

12. What are the characters of chemical changes?

13. What do you consider as the chief difference between a physical and a chemical change?

14. Describe any two experiments which you consider will illustrate the difference between physical and chemical changes.

15. Give some common examples of chemical changes from everyday life.

16. Give six examples of common physical changes, and state why you consider them as belonging to this class, and explain why you do not consider them as chemical, in each case.

17. What kind of change does a piece of iron undergo when it rusts? State fully the reasons for your answer.

18. Point out the difference between a piece of rusted iron and a piece of magnetised iron.

19. When water changes from the liquid state into solid ice or into vapour, what kind of changes does it undergo? Give your reasons for the answer you give.

20. What do you understand by the term "energy"?

21. Describe fully an experiment by which you would demonstrate the nature of a chemical change.

22. What is chemical affinity?

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Preparation for Second Lesson.

APPARATUS AND MATERIALS REQUIRED.

Dinner plate, solid sodic hydrate, crucible and lid, magnesium ribbon, piece of candle, lime-water, two bell-jars, lamp-glasses, scales, wire gauze, glass rod, tile, phosphorus, pith-ball, ebonite, sugar, chlorate of potash, mortar and pestle, mercuric oxide, test tube, cedar splint, flowers of sulphur, copper turnings, silver nitrate, apparatus as shown in Figs. 7 to 11, iron filings, tartaric acid, soda carbonate, sulphuric acid, quicklime, strong ammonia, hydrochloric acid, pipeclay triangle, crucible tongs, funnel, and corks.

EXPERIMENTS TO BE PERFORMED.

Show that when common salt is produced by action of hydrochloric acid upon caustic soda no change in weight occurs.

Burn some magnesium wire and show gain in weight.

Show that carbon, carbonic acid gas, and water are given off by a burning candle.

Show that the products yielded by the combustion of a candle really weigh more than the candle.

Prove that a magnet will act upon iron through space.

Show an electrified body attracting bodies from a distance.

Prove that a hot substance will ignite phosphorus.

Show that chemical changes take place when bodies are in contact, by touching chlorate of potash and sugar with sulphuric acid.

Demonstrate that sulphur and copper combine when heated.

Decompose mercuric oxide by *heating* it.

Show that *light* promotes chemical action by exposing filter paper soaked in nitrate of silver to a bright light.

Prove decomposition of water by *electricity*.

Fuse some sulphur with iron filings, show new body formed.

Prove that a chemical change is promoted by *solution* by adding a solution of carbonate of soda to tartaric acid.

Strike a mixture of chlorate of potash and sulphur with a hammer to show action of *percussion*.

Show action of unlike bodies, iodine and phosphorus, upon each other.

Prove that heat is developed by chemical action of cold water upon cold quicklime.

Show that chemical changes lead to marked alteration in physical properties by action of hydrochloric acid gas upon ammonia gas.

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CHAPTER II.

THE SCOPE OF CHEMICAL STUDY.

Objects and Scope of Chemical Study—Indestructibility of Matter—Chemical Affinity, and how it is promoted by Heat, Light, Magnetism, Electricity, Fusion, Percussion, and Solution — Chemical Action most marked between Unlike Kinds of Matter—Chemical Action leads to Alteration of Special Physical Properties.

We have seen in our last lesson that **chemistry investigates the composition of all kinds of matter,** and aims at discovering the laws which regulate the changes in its constitution. This knowledge has been gained by careful experiment, accurate observation, and exact reasoning; and those who have obtained this knowledge, and by their labours increase it, are called **chemists.** All that we know of matter has been learnt by this method of handling, examination, and observation. Chemistry is therefore termed an **experimental science,** and it is by experiment only that the student may hope to gain some real knowledge of his subject. Bear in mind that **every experiment is really a question put to nature,** to which she always gives a truthful answer, if we only have the training and intelligence to translate the same. We must always be certain what it is we wish to find out, and then

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put our question in a direct form, and be ever watchful for the answer, from which we may afterwards draw our conclusions.

Indestructibility of Matter.—From the standpoint of the student of chemistry the most important property of matter is that of indestructibility. This great fact is really the corner-stone and foundation of the whole science. Our first experiments in this lesson must therefore be designed to illustrate the great truth, upon which the whole of our future study rests, that **matter cannot be destroyed**, although it may undergo an infinite number of changes in form and condition. The experiments chosen are only taken as types of hundreds, and the student will himself be able, as his knowledge of the subject increases, to add others to the list of examples.

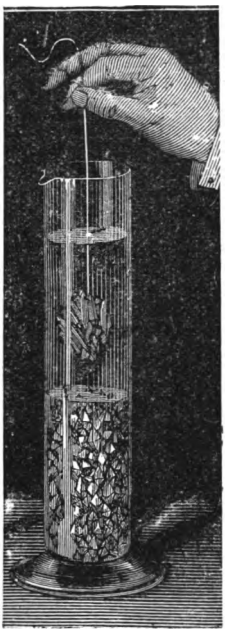

Fig. 4.—Crystals of common salt produced by the action of hydrochloric acid upon sodic hydrate.

Experiment 6.—Place in one glass tube some *hydrochloric acid*, and in another some strong solution of *sodic hydrate*; put the tubes side by side on one pan of a pair of scales, and balance them with weights. Then add the caustic soda drop by drop to the hydrochloric acid, constantly stirring the mixture with a glass rod. After noting that there is no change of weight, allow the liquid to stand in a cool place for some time, or pour a portion into a dinner plate.

Observe that the resulting liquid produced by mixing the two bodies is neither soapy, like caustic soda, nor acid, like the hydrochloric acid; and as it cools white solid crystals are formed, which have the taste of common

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salt. The formation of the crystals is promoted by pouring the liquid into a dinner plate or watch glass.

We learn from this that these two bodies have undergone a real chemical change, for a new substance, common salt, quite unlike each of them, has been produced, yet the *total quantity of matter is the same* as indicated by the balance. There has been no loss or gain of matter, although both the caustic soda and hydrochloric acid have changed their form, and given rise to quite a new kind of matter. This fact is usually stated in the following form: *on the chemical union of two or more substances their original weight remains unchanged.*

The student may here perhaps be inclined to ask the questions: Does no loss of matter occur when bodies are burnt? Are not materials destroyed by fire? The answers are: No loss of matter takes place, and although the forms of substances are destroyed by fire, the matter of which they consist only changes its form, often from the visible into the invisible. Matter itself is no more destroyed during burning than the sugar is, which is dropped into water and there disappears; the water becomes sweet, and if we boil off all the water, we may recover the whole of our sugar.

The two following experiments will illustrate the fact that the matter is not destroyed when a body is burnt. In the first case a solid substance has to be burnt, and a solid but quite different material remains; and in the second case a solid body is to be employed, which is entirely changed into invisible gases during combustion.

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Experiment 7.—(*a*) Place in a small crucible a piece of *magnesium wire*, about three inches in length, twisted into a loose spiral; cover the crucible with its lid, and weigh the whole.

(*b*) Place the crucible with its contents on a clay triangle fixed on a tripod stand, as shown in Fig. 5, and heat it, first gently and then strongly, by means of a Bunsen flame; at the same time raise the lid, but hold it quite near to the crucible, so that no fumes may escape from the burning magnesium, but condense on the cool lid.

(*c*) When the ribbon is burnt allow the crucible to cool, and then weigh the remaining ash with crucible and lid.

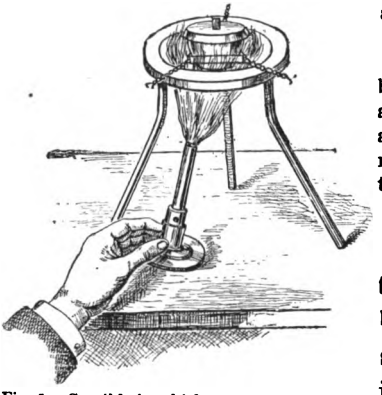

Fig. 5.—Crucible in which magnesia is produced by burning magnesium.

Observe that the magnesium burns and leaves a new substance, a fine white earthy ash of *magnesia*, as in experiment 4, and that this residue really weighs more than the ribbon originally weighed.

We learn from this that the matter which has been burnt is not destroyed, but merely changes its form from a bright lustrous metal into a dull white powder. And we shall be able to prove in later lessons that the increase in the weight of the matter after burning is due to the fact that the metal, like all other bodies which burn in air, combines with some of the gaseous matter of the air called oxygen. This taking up of oxygen is the reason why the ash really weighs more than the metallic ribbon from which it is produced.

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The original substance was magnesium, the resultant body is a compound of magnesium and oxygen called magnesia.

Experiment 8.—(*a*) Place a piece of *burning candle*, about one inch in length, on a glass plate.

(*b*) After the candle has been burning for a few seconds, take a clean piece of white paper or cold porcelain, and draw it backwards and forwards two or three times just above the top of the flame.

(*c*) Place a clean, cold, dry, glass bell-jar over the burning candle.

(*d*) Next pour a little clear lime-water into a large wide-mouthed bottle as shown in Fig. 7. Shake it up and observe that it undergoes little or no change. Pour the lime-water away and fix a burning candle in the bottle, close the mouth with a cork, through which a narrow glass tube passes with a funnel at the top. The flame of the candle quickly dies out.

Fig 6.—Water produced by a burning candle.

(*e*) When the candle has ceased burning, pour some clear lime-water into the funnel, and observe that, on shaking the bottle, the lime-water loses its clearness and becomes milky-white in appearance.

Observe that the burning candle gives off carbon which is deposited upon the cold paper or porcelain, and that it yields steam which is deposited as water upon the cold glass. Also notice that the air in the bottle before the candle is burnt does not change the lime-water; but after the candle has burnt in the bottle for a short time, the air is so changed that it at once makes the lime-water turbid.

We learn from these experiments the following facts:—A candle in burning gives off free carbon. It will not continue to burn in a limited supply of air, and when it burns, water is produced and a gas is liberated which has the power of turning lime-water milky. This last body yielded by the burning candle is termed *carbonic acid gas*.

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If we consider these experiments for a few minutes we shall see that heat produces first a physical change in the solid fat of the candle—in fact, it changes it into liquid fat, which on cooling will pass back again into unaltered solid fat. The liquid fat rises up the wick, and by the greater heat is converted into gases which there combine with some oxygen of the surrounding air, and so produce the water and carbonic acid gas. The candle will only continue to burn so long as it can obtain some free oxygen gas, and when the supply of that body has been used up the flame dies out.

Fig. 7.—Lime-water experiment with candle.

Now if the matter of which the candle consists does really combine with oxygen of air to form the water and carbonic acid gas, which we have proved are produced, we should expect to find the total weight of all the matter yielded by the combustion of the candle greater than the weight of the candle itself. The increase would therefore represent the quantity of oxygen taken up from the air by the burning matter of the candle.

Experiment 9.—(*a*) Fit up a *lamp-glass*, as shown in Fig. 8, with a perforated cork below, upon which a short piece of thin candle is attached, and above place a loose plug of wire gauze containing some fragments of *quicklime* or *caustic soda*, which bodies will serve as a net to catch all the products of combustion.

(*b*) Suspend the lamp-glass when fitted up from one arm of a balance, as shown in Fig. 8.

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(*c*) Carefully counterpoise the apparatus by placing small shot in one pan.

(*d*) Next withdraw the cork, ignite the candle and quickly replace it in the tube, fixing the cork so that air may be drawn in through the holes below.

Observe that as the candle burns and gradually decreases in size, the scale descends and more shot must be added to restore the balance.

We learn from this that the invisible bodies which are

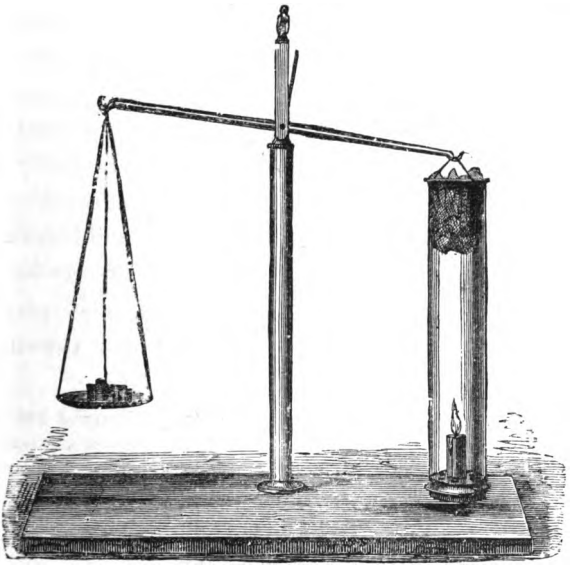

Fig. 8.—Apparatus to show increase of weight during combustion of a candle.

produced when a candle burns really weigh more than the matter of the candle.

Thus, then, we have again seen that matter in undergoing a marked chemical change is not really destroyed.

We do not *know* that the indestructibility of matter is a universal truth, but it has been proved to be true in very many cases ; and as chemists have discovered no exception, *we believe it to be a universal truth.*

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CHEMICAL AFFINITY, we have seen in the last lesson, is the name given to the form of energy which gives rise to those changes of matter which we have learned to distinguish as chemical changes. This force **acts only when bodies are in close contact,** and in this respect it differs from the ordinary physical forces which act upon bodies across intervening space. It is in *all* cases necessary that bodies taking part in a chemical change should be in actual contact.

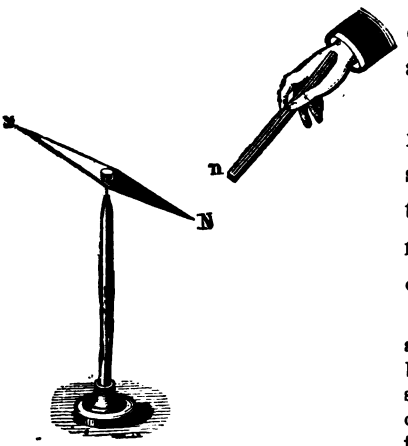

Fig. 9.—Magnetic needle attracted and repelled through a distance by a magnet.

The following experiments should assist the student in remembering the fact that **physical forces act through a distance.**

Experiment 10.—Bring a magnet or magnetised knife blade (experiment 2) near to a small magnetic needle, or ordinary sewing needle, suspended from its middle by means of a little sealing-wax and thread; or place a magnet beneath a piece of thin cardboard upon which some iron filings have been scattered, and give the cardboard a few taps.

Observe that the needle is attracted by the magnet through a distance, and that the iron filings arrange themselves in curves under the influence of the magnet, although they are separated from it by the cardboard.

We learn from this that the *force of magnetism will act upon matter through space.*

Experiment 11.—Suspend a small *pith* or *cork-ball* by means of a piece of silk thread, as shown in Fig. 3. Bring near to the pith-ball an electrified rod of ebonite.

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Observe that the pith-ball is first attracted and then repelled by the electrified rod.

We learn from this that the *force of electricity like that of magnetism will act upon matter through space.*

Experiment 12.—Place a small piece of *phosphorus* on a tile and bring near to it a hot glass rod or poker. Phosphorus is a very inflammable substance, and must, therefore, be kept under water; it should *not be handled by the student*, but small fragments should be distributed by the teacher, when necessary.

Observe that the phosphorus is melted at the surface by the heat, and takes fire, although the hot body does not touch it.

We learn from this that the *force of heat will act upon matter through a distance*, and in this respect we see it behaves like magnetism and electricity.

Experiment 13.—Mix and grind thoroughly some powdered *loaf sugar* and a small quantity of *chlorate of potash*. Place a little of the powder on a tile and bring near the mixture a glass rod which has been dipped in strong *sulphuric acid*, and next touch the mixture with the moistened rod.

Observe that no change takes place so long as the rod is only **near** to the powder, but directly the sulphuric acid is **brought into contact** with the mixture it bursts into flame, and undergoes a chemical change.

We learn from this that chemical force only acts when bodies are in **actual contact**; for whilst the sulphuric acid was only held **near** the mixture no change occurred, but directly it was brought into **actual contact** a marked chemical change took place.

CHEMICAL CHANGES are promoted by all those physical means which tend to promote the contact of matter. We have seen in last lesson that the force of heat tends to reduce a solid to a liquid. It therefore tends also to make easier that more intimate contact of the particles of matter which is necessary before changes of composition can be brought about by chemical affinity.

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Such means as *solution*, *grinding*, and *fusion* are also often employed to facilitate chemical action.

Experiment 14.—Place some *flowers of sulphur* in a small flask, and then half fill the flask with *copper turnings* and weigh the whole. Now heat the upper part of the flask in the neighbourhood of the copper, and then place the flask on a tripod, as shown in Fig. 10, and heat the sulphur. Allow the flask to cool, and weigh again.

Observe that the sulphur melts—boils and gives off vapour—and that the heated copper burns in the vapour of the sulphur, and that a blue-black solid body remains. The resulting substance weighs exactly the same as the two kinds of matter we started our experiment with, but it is quite unlike—brittle yellow sulphur and red metallic copper.

Fig. 10.—Copper turnings and sulphur heated in a flask.

We learn from this that a *chemical change has been facilitated by the action of heat*, for the two cold bodies, sulphur and copper, may be kept in contact for any length of time and no chemical change will take place. But here a new kind of matter has been produced by the direct chemical union of two bodies without any loss or gain in weight. This method of producing a new kind of matter by causing bodies to **combine chemically** is termed **synthesis.** This is the name always used by chemists for the processes employed in building up new bodies by chemical means. In this case it is *sulphide of copper*, which has been formed from sulphur and copper.

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Experiment 15.—Introduce a small quantity of red mercuric oxide into a clean, hard glass tube, and apply a smouldering cedar splint to the open end of the tube. Now heat the portion of the tube containing the red oxide, and again bring a glowing splint to the mouth of the tube.

Observe that the cold mercuric oxide is a bright red powder, and that the glowing cedar splint dies out when first brought to the mouth of the tube. After heating, the powder blackens, the clean tube becomes coated with a bright metallic deposit of mercury, and the glowing splint bursts into flame and burns brightly.

We learn from this that a chemical change has been promoted by the action of heat, for the bodies produced —one bright metallic mercury, and the other an invisible gas, which caused the smouldering splint to burst into flame—are altogether unlike the mercuric oxide with which we started our experiment.

In the above experiment one form of matter has yielded two other substances. This kind of chemical action is due to decomposition, and whenever bodies are thus reduced to a simpler form or broken up chemically, we speak of the action as **analysis.**

In the last two experiments we have seen that chemical changes may be promoted by the action of heat, and we may next learn that **light favours chemical change.**

Experiment 16.—Pour a strong solution of *nitrate of silver* on two strips of filter paper, expose one piece to the sunlight for a few minutes, and place the other in a dark drawer.

Observe that the paper exposed to the light slowly darkens, and finally becomes almost black, whilst that which is protected from the light remains quite white.

We learn from this that a chemical change has been promoted by the action of light, for it may be proved that the alteration in colour is due to a chemical change.

The art of photography depends very largely upon the fact that chemical changes are promoted by the action of

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light. Some of the most characteristic chemical operations which take place on a large scale in nature are due to the action of light, and among these perhaps the most important and most widely known is the absorption of carbonic acid gas and elimination of oxygen by plants.

Experiment 17.—Fit up a jar, *f*, as shown in Fig. 11, with a cork provided with two strips of platinum, *n*, *m*, which terminate above in binding screws, *a*, *b*. Completely fill the jar with water made slightly acid by the addition of sulphuric acid, and pass an electric current through the platinum wires *a*, *b*. After a few seconds connect the glass tube *c*, which passes through the centre of the cork by means of an indiarubber delivery tube, *r*, with a small basin, *g*, containing soap lather.

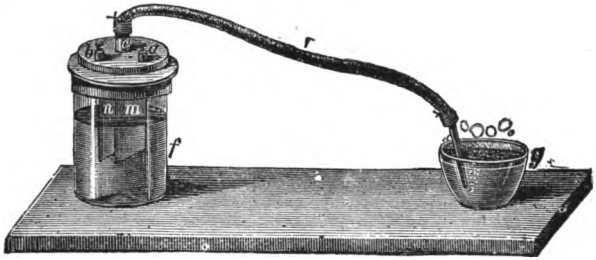

Fig 11.—Apparatus for producing a chemical change in water by the aid of electricity.

Observe that after a short time a gas escapes from the delivery tube which blows up small soap-bubbles; these on the application of a burning taper explode; the gas, therefore, *cannot be steam.*

We learn from this that a chemical change has been brought about by the passage of the current of **electricity** through the water. We shall see in a later experiment that the explosive gas thus produced is a mixture of the two gases oxygen and hydrogen.

We have seen that heat promotes the fusion of bodies, and thus tends to bring the particles nearer together; for this reason many substances when fused exhibit a greater tendency to undergo chemical change.

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Experiment 18.—Mix thoroughly and grind together some *iron filings* and *flowers of sulphur*, and heat the mixture in a crucible.

Observe that before the mixture is heated the iron, although intimately mixed with the sulphur, may be withdrawn by means of a magnet. During heating the mass melts or fuses, and finally a grey substance remains, very unlike either iron or sulphur; from this body iron cannot be extracted by means of a magnet.

We learn from this that a change due to the chemical union of the iron and sulphur has been **promoted by fusion.**

The close contact of matter may be promoted by dissolving the substances experimented upon.

Experiment 19.—Prepare a solution of *carbonate of soda* by shaking a small quantity of solid carbonate of soda with water. By shaking a crystal of *tartaric acid* well with water prepare, in a second tube, a strong solution of tartaric acid. Add some of the solution of tartaric acid to the solution of carbonate of soda.

Observe that violent effervescence occurs, and that a gas escapes which if passed into lime-water renders it milky. The resulting liquid acquires a saline taste.

We learn from this that a chemical change has taken place when the two **solutions** are brought in contact.

Experiment 20.—Now mix and grind well together some *solid carbonate of soda* and solid *tartaric acid.*

Observe that so long as the mixture is kept dry no chemical change is produced, but that the resulting body has all the properties of both of the substances employed. But on the addition of a little water to the mixture violent effervescence takes place, indicating that then a most decided chemical change occurs.

We learn from this that no chemical change is produced when dry carbonate of soda and tartaric acid are mixed. This experiment considered and compared with the last should assist the student in understanding that **chemical changes may be promoted by solution.**

Experiment 21.—Mix a small quantity of finely-powdered *chlorate of potash* and *flowers of sulphur.* Place a very small portion of the mixture on an iron plate and give it a sharp blow with a hammer.

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Observe that no material change takes place on mixing the bodies, for the chlorate of potash may be separated from the insoluble sulphur by the action of water. On striking the mixture a sharp explosion occurs and a strong-smelling gas escapes.

We learn from this that **chemical change** has been **facilitated by percussion.**

Substances which are most unlike usually exhibit the greatest tendency to act chemically upon each other when they are brought in contact.

Experiment 22.—Place a piece of *phosphorus* about the size of a small pea on a tile, and then drop a few fragments of *iodine* upon it.

Observe that the two bodies are very unlike. The phosphorus is a yellow, waxy, highly-inflammable solid, and the iodine is a lead-grey solid, which stains the fingers brown, and which when heated even does not ignite, but merely changes into a liquid, and then is converted into a beautiful violet vapour; but when these two different kinds of matter come in contact a most marked chemical change takes place.

We learn from this that those forms of matter which are most unlike act with the greatest readiness chemically upon each other.

Chemical action always gives rise to a change in temperature in the matter taking part in the change; in most cases **heat is produced.** For example, in experiments 7, 9, and 14, we saw that so much heat was developed by the chemical action of the bodies which took part in the changes that they gave out light.

Experiment 23.—Place a large piece of freshly-prepared *quicklime* on a dinner plate, make a small hole in one side of it, and place in it a piece of freshly-cut *phosphorus* about as large as a hemp-seed. Now pour carefully some *cold water* on the *cold quicklime*, but do not wet the phosphorus.

Observe that the water disappears, the lime becomes hot and swells, clouds of steam are given off, and the hot lime melts the phosphorus, which afterwards bursts into flame.

We learn from this that these two cold substances, by their chemical action upon each other, **liberate heat.**

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Chemical changes give rise to an alteration in the special physical properties of the bodies taking part in the change. The *specific properties*, such as taste, odour, condition, and colour, are changed. All the foregoing experiments bear out this statement, but one other example will not be out of place here, for the student can only hope to secure a fair foundation of knowledge by continued experiment.

Experiment 24.—(*a*) Pour a few drops of strong *hydrochloric acid* into a gas jar, *B*; shake well and decant the liquid to the acid bottle; now close the jar, which contains hydrochloric acid gas, with a glass plate.

(*b*) Into a second jar, *A*, pour a small quantity of strong *ammonia solution*, shake, and then turn the liquid back into the bottle, and cap the jar with a glass plate. This jar contains invisible, strong-smelling ammonia gas.

(*c*) Place the capped jars mouth to mouth, as shown in Fig. 12, so that the jar of hydrochloric acid gas stands upon that containing ammonia gas; then carefully withdraw the plates, *D*.

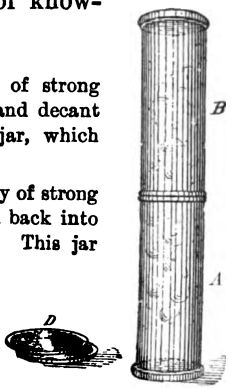

Fig. 12. — Action of hydrochloric acid gas upon ammonia gas to form solid chloride of ammonia.

Observe that directly the gases come in contact they combine to form a cloud of fine white solid substance, which falls to the bottom of the lower jar in white flakes like very fine snow.

We learn from this that two very different *gases*—one fuming acrid hydrochloric acid, the other strong-smelling pungent invisible ammonia—form together a new body, which is a white *solid* without odour, but with a taste not unlike that of common salt. Hydrochloric acid gas and ammonia chemically unite to form the new substance **ammonium chloride.**

As we progress with our study we shall learn that such great changes as these are characteristic of chemical action.

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SUMMARY TO CHAPTER II.

CHEMISTRY - - Investigates the changes which take place in the composition of matter; it aims at the discovery of the causes of such changes and at an exact statement of the rules regulating such.
Exact statements of groups of facts are *natural laws.*
An *experimental science* of which we learn something by the handling and examination of matter.
First law of chemistry—*indestructibility of matter.*

PHYSICAL FORCES - - - Such as heat, magnetism, electricity act upon matter from a distance. Thus phosphorus is ignited by a hot rod, iron is attracted by a magnet from a distance, and light substances are attracted and repelled by an electrified body.
The changes produced by these forces are temporary in character, and are not attended by alterations in the weight of the matter taking part in the change.

CHEMICAL AFFINITY - - - - 1. Gives rise to changes only when one kind of matter is in *actual contact* with other matter.
2. Promoted by *heat, light, grinding, solution, fusion, electricity.*
3. Acts most readily between forms of matter which are most unlike.
4. Heat is usually developed through chemical action, although in some few cases heat is absorbed during chemical change.
5. Leads to changes in the special physical characters of matter. So that forms of matter produced by the chemical union of two or more bodies possess properties usually entirely unlike those which were characteristic of the components.

ANALYSIS - - - The breaking-up of a substance into two or more different kinds of bodies.

SYNTHESIS - - The building-up of one substance from two or more different kinds of matter.

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QUESTIONS ON CHAPTER II.

1. What is chemistry? Why is chemistry called an experimental science?

2. "Matter is indestructible." How would you prove this statement to be true?

3. What do you understand by the term "chemical affinity"?

4. What is the great difference between "chemical affinity" and the so-called "physical forces"? Give examples of this difference.

5. Show how you would prove that physical forces will act upon matter through a distance.

6. Why do the substances produced from a body burnt in air weigh really more than the substance which is burnt?

7. What bodies are produced when a candle is burnt in air?

8. Give an example of a chemical change, and describe fully an experiment which proves that chemical affinity acts only when bodies are in actual contact.

9. How would you prove that chemical affinity is greatest between kinds of matter most unlike?

10. Describe any two experiments by which you would prove that chemical action is promoted by (*a*) heat (*b*) electricity.

11. What do you understand by the term "chemical analysis"?

12. Explain what you understand by synthesis, and give an experimental example.

13. Show how you would prove that light facilitates chemical action.

14. Give three examples of chemical changes from everyday life.

15. Why do you class the changes which occur when bodies are burnt in air as chemical?

16. How would you prove that bodies undergo a change in physical properties when acted upon by the force of chemical affinity?

17. Describe two experiments by which you would demonstrate that heat is usually evolved when chemical changes take place.

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Preparation for Third Lesson.

APPARATUS AND MATERIALS REQUIRED.

Platinum wire, sulphur, copper turnings, red mercuric oxide, sulphide of copper from experiment 14, iron filings, nitre, charcoal, a mortar and pestle, specimens of metals, non-metals, alloys, and amalgams.

EXPERIMENTS TO BE PERFORMED.

Repeat experiments 7, 14, and 15.

Show that nitre and charcoal may be mechanically mixed, and that the components of the mixture may be identified and separated. Point out the great difference which exists between the products of their chemical combination and the original components of the mixture.

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CHAPTER III.

CHEMICAL CLASSIFICATION OF MATTER.

Chemist's Division of Matter—Elements and Non-Elements—Chemical Compounds and Mechanical Mixtures—Metallic and Non-Metallic Elements.

CLASSIFICATION OF MATTER.—We have seen in our first lesson how we are surrounded by an infinite variety of different kinds of substances, and we have learnt that all may, for convenience of study, be arranged according to their physical condition as **solids** and **fluids.** This classification depends upon the physical state in which the body exists, and has nothing to do with the question as to what it is made up of. Chemistry, we have seen, inquires into the deeper question as to what all kinds of matter are made up of; it becomes necessary therefore to adopt some other method of grouping bodies. Students of chemistry have adopted a simple division of all kinds of matter into two classes: in one they place all those bodies from which they can by any means extract two or more different kinds of matter, and in the other class they include all those substances from which they cannot extract any new kind of matter by any known means. These two groups are known as **non-elements** and **elements.** In arranging substances according to this method of grouping

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we disregard the physical state in which the body exists, and consider only the facts mentioned above; for example, solid ice, liquid water, and the gases, steam, and air are all non-elements, whilst solid sulphur, liquid quicksilver, and gaseous oxygen are elements.

ELEMENTARY MATTER includes all that which cannot be reduced to a simpler form; thus we cannot extract any other body from the platinum employed in experiment 1, nor can we withdraw any new body out of sulphur; although we may make it undergo many changes in condition, yet it still remains sulphur, and nothing else. For these reasons both platinum and sulphur are considered *elements.* In experiment 15 we saw that, by heating mercuric oxide, it was possible to extract two very different bodies from it, one a metal, the other a gas; mercuric oxide is therefore classed as a *non-element.*

NON-ELEMENTARY MATTER is that kind of stuff from which different bodies may be extracted. In certain cases the distinct kinds of matter present may be isolated by simple mechanical means; thus sand shaken with water when allowed to stand soon separates, because the sand is heavier than the water; in such a case the bodies are said to be *mechanically mixed.* Other substances may be made to yield their constituents only with difficulty, for the substances of which they are made up are held together by the much stronger force of chemical affinity; we had examples of such in the common salt of experiment 6, and the magnesia of experiment 7. The components of such non-elementary forms of matter are said to be *chemically combined.*

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The two distinct forms of non-elementary stuff are therefore known as **chemical compounds** and **mechanical mixtures.** In experiment 14 we learnt by a process of synthesis that sulphide of copper must be a chemical compound, for we built up this body from the two different substances sulphur and copper. And again in experiment 15 we proved that red oxide of mercury must belong to the same class of substances, for we resolved it by analysis into the two dissimilar bodies oxygen gas and metallic mercury. Mechanical mixtures are produced by the more or less intimate mixture of different kinds of matter; in such the component bodies retain their original distinctive properties, thus in experiment 18 iron and sulphur when ground together yielded a mechanical mixture in which the iron retained its distinctive magnetic properties, for we saw it could be removed by means of a magnet; if a small portion of the mixture be thrown into water the heavier iron filings at once sink, whilst the sulphur floats. In experiment 21 the white powder produced by mixing chlorate of potash and sugar retained the saltness of the chlorate of potash and the sweetness of the sugar. How different were the bodies which were produced when chemical affinity caused the two kinds of matter to act upon each other and combine to form new substances!

Experiment 25.—(*a*) Grind together a small quantity of dry nitre and dry charcoal by means of a pestle and mortar.

(*b*) Place a portion of the mixture in water and well shake it, and then pour it into a filter as shown in Fig. 17, and evaporate a portion of the liquid which passes through the filter paper to dryness.

(*c*) Put another portion of the dry mixture on a tile or iron plate, and touch it with a burning match.

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Observe that the nitre is a whitish salt body, and the charcoal a black tasteless powder; observe also that the finely ground mixture is grey, and partakes of the properties of both charcoal and nitre. When a portion is thrown into water the carbon floats, and after filtering, the clear water is salt to the taste, and yields nitre, whilst charcoal remains on the filter paper.

On the application of heat the mixture burns, and the components undergo a decided chemical change.

We learn from this that nitre and charcoal ground together only form a **mechanical mixture,** no matter how fine the powder may be. The individual particles of such a mixture may be recognised with ease by means of a magnifying-glass, and the substances can be separated by a simple physical experiment. When the constituents of mixtures combine they form entirely new substances.

CHEMICAL ELEMENTS are variously associated, combined or mixed to form all the many different kinds of materials known to us. At present chemists have discovered about **seventy elements** which they regard as ultimate kinds of matter. Some few only are met with free or uncombined in nature, but most are chemically united in groups of two or more. A few are very abundant and widely distributed, whilst many others are exceedingly rare. It is only those which are most abundant and important which we shall study in these lessons.

The chemical elements are usually subdivided into two groups—the metals and the non-metals.

THE METALLIC ELEMENTS, with the exception of mercury, are all solids at ordinary temperatures. They possess, as a rule, a certain well-marked brightness called **metallic lustre;** they are good conductors of heat and electricity; they can usually be drawn into wire and beaten into thin plates. Hence they are said to be both

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ductile and malleable, although some few are quite brittle. As a class they are heavier than water, but here, again, there are a few exceptions. The metals *sodium* and *potassium*, we shall see, will both float on water.

More Important Chemical Elements.

Name.	Symbol.	State.	Name.	Symbol.	State.
Aluminium ...	*Al*	Solid	*Lead*	*Pb*	Solid
Arsenic	As	„	Magnesium ...	Mg	„
Barium	Ba	„	Manganese ...	Mn	„
Bismuth... ...	Bi	„	*Mercury*	*Hg*	Liquid
Boron ...	**B**	„	**Nitrogen** ...	**N**	Gas
Bromine ...	**Br**	Liquid	**Oxygen** ...	**O**	„
Calcium... ...	*Ca*	Solid	**Phosphorus**..	**P**	Solid
Carbon ...	**C**	„	Platinum... ...	Pt	„
Chlorine ...	**Cl**	Gas	Potassium ...	K	„
Copper	*Cu*	Solid	**Silicon** ...	**Si**	„
Fluorine ...	**F**	Gas	Silver	Ag	„
Gold	Au	Solid	*Sodium*	*Na*	„
Hydrogen ..	**H**	Gas	**Sulphur** ...	**S**	„
Iodine ...	**I**	Solid	*Tin*	*Sn*	„
Iron	*Fe*	„	Zinc	Zn	„

The names of the non-metallic elements are printed in thick type, thus: **Oxygen O**, and those of the metallic elements which are dealt with in this work are given in italics, thus: *Aluminium Al.* A complete list, including all the rarer elements, will be found in the Appendix.

The **most abundant metals** in the earth are *aluminium*, *calcium*, *sodium*, *magnesium*, and *potassium*, none of which are found in the free condition in nature, but always combined with one or more other elements. In addition to these, *iron*, *lead*, *zinc*, *tin*, *copper*, *nickel*, *mercury*, *silver*, and *gold* are of importance, because of their great value in the arts and manufactures. Very intimate mixtures of two

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or more metals produced by fusion, like *brass*, *bronze*, and *German silver*, are known as **alloys.** When mercury is one of the components of the mixture, the resulting body is termed an **amalgam.**

THE NON-METALLIC ELEMENTS in a more or less marked degree exhibit the absence of those properties by which we may distinguish the metals. Four bodies of this class—namely, *oxygen*, *hydrogen*, *nitrogen*, and *chlorine*—exist under ordinary conditions in the gaseous state, and *fluorine* has recently been isolated as a gas. One, *bromine*, is a liquid ; all the rest are solids.

The most abundant non-metallic elements, making up the materials of the solid earth, are *silicon*, *oxygen*, *sulphur*, *chlorine*, and *hydrogen* ; whilst *nitrogen* makes up about four-fifths of the whole atmosphere. In substances of vegetable and animal origin, the most abundant elements are *carbon*, *oxygen*, *hydrogen*, and *nitrogen*.

Although the possession of certain distinctive physical properties has been adopted as a means of classifying the elements as metals and non-metals, yet it is well to bear in mind that there is no very hard and fast line of demarcation ; for the properties mentioned are possessed in very varying degrees, and one group really merges into the other, certain of the elements standing as it were on the border line. Thus the non-metals, iodine and arsenic, have considerable lustre, and the latter, for strictly chemical reasons, is often regarded as a metal. The metals lithium, sodium, and potassium are so light that they will float on water ; and the non-metals bromine, iodine, and arsenic are from three to five times as heavy as water.

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SUMMARY TO CHAPTER III.

ELEMENTARY, SIMPLE SUBSTANCES — Bodies which cannot be reduced to a simpler form by any at present known means. *An Element* is a body from which we *cannot* extract essentially different kinds of matter.

NON-ELEMENTARY — Bodies from which we can extract two or more essentially different kinds of matter.

ELEMENTS

Metals: About *fifty-five* are known, all of which are *solids* save mercury; they are all more or less *ductile, malleable*, good *conductors of heat* and *electricity*; and they usually have a *high specific gravity*. They form, with *oxygen*, compounds called *oxides*, all of which possess the *power of neutralising acids*.

Non-Metals: About *fifteen* are known. One exists in the liquid state, four are gaseous, and the remainder are solids. Those which exist in the solid state are usually *brittle* and *bad conductors of heat* and *electricity*. They form *oxides* which all *dissolve* more or less readily in water *to form acids*.

NON-ELEMENTS

Chemical Compounds: Bodies consisting of two or more elements *united* by the force of *chemical affinity*.

Mechanical Mixtures: Bodies consisting of two or more substances, *elementary or compound*, more or less intimately mixed. *Alloys* and *amalgams* rank under this head rather than as chemical compounds.

QUESTIONS ON CHAPTER III.

1. Explain the terms "element" and "non-element."

2. What is the difference between a mechanical mixture and a chemical compound?

3. Classify the following substances into elements and compounds: — Steam, ice, sulphur, ammonia, common salt, marble, and carbonic acid gas.

4. How would you produce a mechanical mixture? How would you prove that the substance which you had prepared was only a mechanical mixture?

5. How would you describe a mechanical mixture?

6. What is a chemical compound? Give four examples and state fully how you would prepare any one of them.

7. What are the more important properties which enable you to decide whether an element is classed as a metal or non-metal?

8. Classify the following elements into metals and non-metals:—Aluminium, calcium, carbon, chlorine, copper, fluorine, hydrogen, iodine, iron, lead, manganese, mercury, nitrogen, oxygen, phosphorus, potassium, silicon, silver, sodium, sulphur, zinc.

9. Describe minutely what occurs when hydrochloric acid gas and ammonia gas are mixed together. Give a sketch of the apparatus which you would employ to show this experiment.

10. What is the difference between a mixture of two elements and a compound of two elements?

11. Classify the chief elementary substances into metals and non-metals.

12. Divide the following substances into elements and compounds:—Glass, lime, iron, gold, water, ammonia, flint, chlorine, and nitrogen.

13. By what properties is a metal distinguished from a non-metal? Arrange the following elements under the above heads:—

Chlorine	Nitrogen	Iron
Potassium	Sodium	Zinc
Phosphorus	Arsenic	Lead
Sulphur	Hydrogen	

14. Name the metallic elements which are most abundant in the earth.

15. Give a list of the most abundant non-metallic elements present in the solid earth.

16. What are the most abundant elements in living things?

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Preparation for Fourth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of washing soda, alum, sugar, salt, balance, weights, beakers and glass rods, sandbath, evaporating dish, tripod stand, watch and clock glasses, specimens of any large crystals or models in wood or glass of crystals, iodine, nitre, black oxide of manganese, filter fitted up as in Fig. 17, specimens of natural water, shellac, sealing-wax, camphor, indiarubber, fat, methylated spirit, ether, carbon-disulphide, chloroform, olive oil, mercury, long glass tube, apparatus fitted up as in Fig. 19, specimens of hartshorn and spirits of salt.

EXPERIMENTS TO BE PERFORMED.

"Suspend a piece of white sugar, on a tray of perforated zinc or wire gauze, by a thread in a glass vessel containing water.

Dissolve salt in water.

Show on a balance that sugar or salt and water when separate and when dissolved weigh the same.

Show that salt is obtained from the solution by evaporation.

Saturate water with nitre, and show that the solubility is increased by increase of temperature.

Demonstrate the formation of crystals.

Illustrate the removal of substances in suspension, and the non-removal of substances in solution by adding water to a mixture of black oxide of manganese and sugar, and then filtering.

Show by evaporation the solid matter dissolved in a sample of pump, or river, or spring water, and explain the method for its quantitative determination.

Show the like solvent action with other liquids, as camphor or shellac in alcohol, and fat or caoutchouc in ether.

Compare the result of admixture of spirit or oil of vitriol with water with that of oil or mercury with water.

Heat ordinary water and collect the expelled air."—*Science Department's Syllabus.*

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CHAPTER IV.

SOLUTIONS AND SOLVENTS.

Solvent Properties of Water—Disappearance of a Solid in Water by Solution—Saturation—Effect of Increase of Temperature on Saturation—Effect of Lowering the Temperature on Saturation—Crystallization—Filtration—Rain, Spring, River, Pond Waters, etc.—Solid Matter in Different Waters: how Estimated—Loch Katrine Water—Thames Water—Sea Water—Hard and Soft Waters—Mineral Waters—Similar Solvent Action of other Liquids—Solution of one Liquid in another—Liquids Insoluble in one another—Solution of Gases in Water and other Liquids—The Effect of Heat on the Quantity of Gas dissolved by a Liquid.

SOLUTIONS.—Some solid forms of matter when placed in liquids disappear, and at the same time the liquids, in which they melt away, become endowed with new properties. Thus *soda*, *sugar*, *salt*, and *alum* all melt or disappear when placed in water; but the liquid becomes *soapy*, *sweet*, *salt*, or *astringent* to the sense of taste as the case may be. The substances which disappear cannot therefore be lost or destroyed, for they impart quite new characters to the water. This disappearance of certain solids when placed in liquids is due to the fact that the liquid has the power of overcoming the force which holds the particles together; and any means by which we may increase the surface of the solid, as for instance grinding, tends also to facilitate the rate at which it will be taken

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up by the liquid. Solid substances which so behave are said to **dissolve** in the liquid, which is spoken of as the **solvent.** This is perhaps the most important property of water, and many of the changes which have taken place on our earth in the past have been brought about through the solvent action of water.

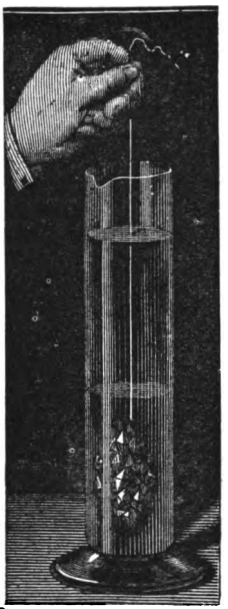

Fig. 18.—Disappearance of a solid in water by solution.

Experiment 26.—Place a small lump of *sugar* upon a piece of wire gauze suspended by threads in a tumbler of water.

Take another portion of sugar and *grind to a fine powder*, and then transfer to a second vessel containing water.

Observe that the sugar *gradually* disappears in the first case and is diffused through the water, which acquires a sweetish taste. Notice that in the second case the sugar rapidly dissolves.

We learn from this that sugar will **dissolve** in water, and that grinding or powdering aids solution. Such an intimate mixture of solid and liquid is termed a **solution**; the water here is the **solvent** or **menstruum,** in which the sugar is said to be **soluble.**

Experiment 27.—On one pan of a pair of scales place a beaker containing about *half a pint of water*, and by the side of it on the same pan place a block of *salt* weighing about 2 oz. Carefully counterpoise the whole by means of weights. Now add the salt to the liquid.

Observe that on the addition of the solid salt to the water no change in weight occurs, nor is there any alteration in weight when the whole of the salt is dissolved.

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We learn from this that no loss or gain of matter takes place when the salt enters into solution, for the total weight of the salt and water when separate and when dissolved is the same. The matter therefore which *disappears* in the water is not lost or destroyed, but merely *added* to the water.

Evaporation.—Bodies held in solution may be recovered from the solvents which hold them, by driving off the liquid as a vapour through the agency of heat, and the vapour may be reconverted into a liquid by withdrawal of heat.

Experiment 28.—Place some of the *solution of salt* obtained in the last experiment in an evaporating dish, D, Fig. 14, on a sand-bath, C, and heat this salt solution contained in the basin by means of a Bunsen burner, A. Take care that the liquid is allowed only to boil gently.

Observe that steam escapes, and that after a time a white deposit makes its appearance on the sides and at the bottom of the basin.

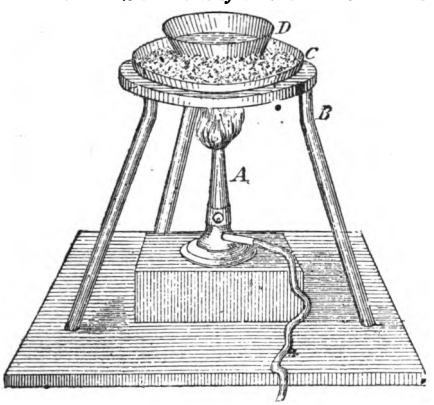

Fig. 14.—Evaporation of a solution of common salt.

We learn from this that salt may be recovered from its solution by the process of evaporation. For by the application of heat the whole of the water is sent off as vapour.

Saturation.—The first portions of a solid disappear most readily in a solvent, and as the quantity of the body increases, the power of the liquid to dissolve it decreases, until finally a limit is reached and the liquid will take up no more. This stage is known as saturation, and when a

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liquid is in this condition, if more of the solid is added it remains undissolved at the bottom of the vessel.

As a rule the solvent power of a liquid may be considerably increased by the application of heat, and a hot saturated solution of any body will therefore contain more of the substance than a cold saturated solution. It follows from this that a hot saturated solution on cooling must throw out of solution some of the solid body which has been dissolved. And since the solvent power of a liquid is usually increased by heat, we see how it is that some substances which are insoluble, or only slightly soluble in cold water, may be dissolved at once by hot water. Bodies which, like flour, sand, clay, and glass, will not melt in water are said to be **insoluble.**

Experiment 29.—Add some powdered nitre to a little cold water in a beaker, constantly stirring; continue this until the solid ceases to disappear. Then place the beaker of cold saturated solution on a sand-bath, and heat it. When hot add more nitre and stir well.

Observe that when the cold water is saturated with nitre some of the solid remains undissolved at the bottom of the vessel, but this dissolves, and the water will take up more nitre on heating.

We learn from this that the solvent power of water increases when the temperature is raised. This is a general rule to which we shall find there are a few exceptions.

CRYSTALLIZATION.—Many solids, as they are thrown out of solution, assume definite geometrical forms called crystals. These vary in size, but are of a fixed form for each different body; they may be *cubes*, or *pyramids*, or *prisms*, etc.

Different solid bodies may be prepared in the crystalline form by various means, thus crystals of nitre, alum, washing soda, and sulphate of copper may be obtained with ease

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from the saturated solutions of these bodies. A weak solution of a substance may be converted into a **concentrated solution** by evaporation, by which means of course liquid is got rid of, and the solution becomes a hot saturated solution, from which crystals of the dissolved solid will be deposited on slowly cooling.

If in experiment 27 the basin containing the solution is removed from the sand-bath just before the liquid has entirely disappeared, minute crystals of salt, which may be examined by means of a magnifying-glass, will form upon the sides of the basin as the solution cools.

Fig. 15.—Crystallization.

Crystals of iodine may be prepared by heating some of the solid in a small evaporating dish, over which a flask of cold water has been placed. The iodine is converted into a liquid and then vaporised, and afterwards condenses upon the cold flask in the crystalline form. When solid bodies are thus converted into vapours which pass at once into solids again on cooling, they are said to undergo **sublimation.**

Experiment 30.—Place some of the *solution* of nitre prepared in experiment 29 on one side to cool.

Observe that as the saturated solution cools some of the solid nitre is thrown out of solution in the crystalline form.

We learn from this that *crystals of nitre* may be obtained by first preparing a hot saturated solution of the body, and then gradually cooling the same.

FILTRATION.—Bodies which are insoluble in water may be mixed with it, and remain **mechanically suspended**

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through the liquid for lengths of time, which vary with the fineness of the particles and the specific gravity of the substance; but such suspended particles sooner or later collect and rise to the top, or fall to the bottom of the liquid. When the particles are very fine, this separation of the insoluble but mixed particles from a liquid may take some considerable time; in order therefore to facilitate this removal, chemists adopt the process of filtration.

A filter is prepared, as shown in Fig. 17, where a glass funnel is placed in a filter stand, and filter paper is doubled first into two and then into four, and then opened and

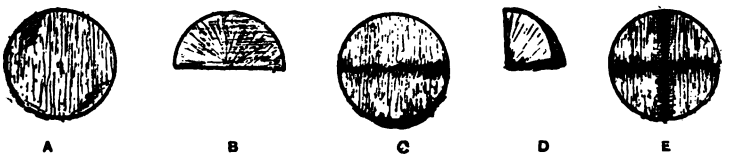

Fig. 16.—Preparation of a filter paper.

fitted into the glass funnel F. Such a filter paper acts as a sieve, for through the pores of the paper the liquid holding a body in solution will pass, but any insoluble substance held in mechanical suspension in the liquid will be deposited upon the filter paper.

Experiment 31.—Mix a small quantity of *sugar* with *black oxide of manganese* and grind the mixture well, and stir it thoroughly with some water in a glass beaker, A. Next pour the whole liquid into a filter, such as shown in Fig. 17, and catch the liquid which passes through in a beaker, B.

Observe that the black oxide of manganese remains on the filter paper; the liquid which passes through is quite clear, but sweet to the taste.

We learn from this that insoluble black oxide of manganese may be removed from the liquid in which it

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is held in suspension by filtration, but that the soluble body, sugar, cannot be taken out of its solvent by such means.

Where the solid substance suspended in water is very heavy it will subside if the liquid is allowed to stand; the clear liquid may then be removed by gently pouring off; this operation is termed **decantation.**

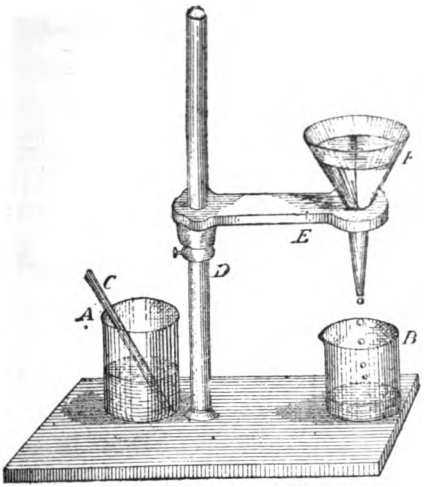

Fig. 17.—Filtration.

NATURAL WATERS differ very much as regards the nature and quantities of substances dissolved in them, but *pure water* is not met with in nature. The most obvious differences between various natural waters are due to the kinds and amounts of soluble materials which they contain. For convenience we usually speak of natural waters according to their modes of occurrence, as sea water, river and lake water, spring water, well water, and rain water.

It is also usual to divide natural waters into two classes,

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according to their action with soap; these are known as hard waters and soft waters.

Rain water.—In districts remote from towns, in the open country and at sea, rain water contains little but dissolved gases derived from the air through which it has passed on its earthward journey; but in the neighbourhood of towns or manufacturing districts many solid and gaseous impurities are washed out of the air by rain.

Well waters are usually soft, but often contain organic matter in solution, which consequently render them more or less dangerous for drinking purposes; they may be rendered safer for use by boiling.

Spring waters are generally clear and sparkling. They rise naturally from rocks, and often have their origin at great depths; in consequence of the thorough filtration such water has undergone, it is usually free from organic and suspended matters. On the other hand, these waters are often charged with saline substances in solution, which they have derived from the rocks through which they have passed. Some spring waters contain so much mineral matter in solution that they are quite unfit for ordinary drinking purposes, although they may be of high medicinal value. Such waters are known as **mineral waters.** *Salts of iron, magnesia, potash, lithia,* and *lime,* and the following acid bodies, *carbonic sulphurous* and *hydrosulphurous,* are often present in such large quantities as to impart decided flavours. Natural mineral waters are named according to the predominating dissolved substance, which imparts some special character, thus :—

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Chalybeate waters contain	salts of iron.		
Saline	„	„	„ magnesia or soda.
Calcareous	„	„	„ lime.
Acidulous	„	„	carbonic acid.

River waters vary very much in character, according to the nature of the district through which they flow. They contain as a rule varying quantities of organic substances derived from cultivated land and drainage, and also substances in solution which are dependent upon the nature of the rocks from which they flow. River waters generally contain materials both in suspension and solution; in England they are as a rule hard, whilst in Wales and Scotland, where the rocks are harder and contain less limy matters, they are more frequently soft, and therefore better adapted for washing purposes. The **Thames water** contains about 23 grains per gallon of dissolved substances, and a variable quantity of organic and inorganic matter held in suspension.

Lake waters contain varying percentages of organic and inorganic substances in solution and suspension similar to those found in river waters, the nature of the water in every case depending upon the character of the rocks of the surrounding district. Thus **Loch Katrine** is in a district surrounded by some of the hardest known rocks, and those upon which air and water can only act very slowly indeed; it is not surprising therefore to find that this water is singularly free from impurities. Extreme cases, which may be compared with it, are those of the Dead Sea and Lake Utah, the waters of which are highly charged with soluble bodies.

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Loch Katrine water	contains about	4	grains	per	gallon	of soluble matter.
Thames water	„	„ 23	„	„	„	„
Sea water	„	„ 2392	„	„	„	„
Lake Utah water	„	„ 12800	„	„	„	„

Sea water far from land, or even quite near land where the rocks are hard, is clear, but strongly saline and somewhat bitter to the taste. It contains about 3½ per cent. of dissolved matter, the most abundant being common salt. In addition to common salt, sea water contains in solution *potassium chloride*, *magnesium chloride*, *magnesium sulphate*, *magnesium bromide*, *calcium bicarbonate*, and *calcium sulphate*. As these bodies are dissolved in sea water they may be recovered from it by evaporation; and as many of them are of value for medicinal purposes, and some are of importance for manufacturing processes, they are extracted on a large scale from brine and from the ashes of seaweeds. Sea water curdles ordinary soap; it cannot therefore be employed for washing purposes; and the salts it contains render it quite unfit for drinking.

Hard waters include all those which are harsh to the sense of touch, and which will not easily form a lather with soap. The hardness is due to the presence in solution in the water of saline substances, which act upon soap and convert it into an insoluble compound, and it is this which forms the scum or curd on the surface of hard water to which soap has been added. Water in which soap lathers only with difficulty usually contains salts of magnesia and lime in solution. Certain hard waters may be made soft by boiling, and others may at least be reduced in hardness by the same means; such are said to be **temporarily**

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hard. This temporary hardness is generally due to the presence of bicarbonate of lime in solution, which body is converted into insoluble carbonate of lime by boiling. For laundry purposes such waters may also be softened by the addition of soda.

Permanently hard waters are those which cannot be softened by boiling; the hardness of such is due to the presence of sulphates of lime or magnesia, or chlorides of magnesium, or nitrates in solution. The hardness of London water is mainly due to the presence of bicarbonate of lime in solution; for this reason this water may be softened to a very considerable extent by boiling, and may therefore be described as temporarily hard. Trent water, on the other hand, and that of many of the rivers of the north of England, mainly owe their hardness to the presence of salts of magnesium, which are not removed by boiling; they are therefore permanently hard.

Soft waters are those which are soft to the touch and do not curdle soap, but easily form a lather with it. Such waters do not contain materials in solution which form insoluble soaps.

Experiment 32.—Place separate portions of tap water, distilled water, and rain water in watch glasses, and evaporate them to dryness.

For lecture purposes it is best to use a large clock glass, and to place it on a wire triangle resting on a beaker of water; the water in the beaker may be boiled, and the steam given off evaporates, slowly, the liquid in the clock glass.

Observe that the tap water leaves a fine white film upon the glass, of matter which was held in solution, whilst the distilled water and rain water leave no residue, and, in fact in the case of distilled water, the glass is as clean after the experiment as before.

We learn from this that the solids dissolved in tap

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water may be recovered by evaporation; and the same is true of all mineral matters held in solution.

If it is required to **determine the proportion of solid matter** held in any specimen of river, lake, or well water, we should take a measured quantity of water and filter it, as in experiment 31, and thus remove the suspended material; the residue on the filter paper must be dried and weighed. A measured portion of the filtered water must then be evaporated to dryness, as in experiments 28 and 32, and the dissolved material thus recovered should be weighed. The two added together give the quantity by weight of solid matter in a given amount of water.

SOLVENT ACTION OF OTHER LIQUIDS.—In the foregoing we have seen that water is a solvent agent, and in nature in consequence of this property it has performed an enormous amount of work in changing the earth in the past, and that work it is continuing to-day. This property makes it also of immense importance in the kingdom of living things, for it is the agent by which solid substances are carried into plant and animal tissues. In all the examples quoted above water has been employed as a type of a solvent liquid, but many solid substances which will not melt in water may be dissolved by other liquids.

Experiments 33, 34, 35, 36, 37.—Place in five test-tubes a small quantity of water, and to one add a little powdered *shellac*, to the second some *sealing-wax*, to the third a little *fat*, to the fourth a small piece of *indiarubber*, and to the fifth a piece of *camphor*.

Observe that these bodies do not dissolve in water, neither on shaking nor on heating.

We learn from this that shellac, sealing-wax, fat, indiarubber, and camphor are insoluble in water.

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Experiments 38, 39, 40, 41, 42.—Place in five other test-tubes pieces of *shellac*, *sealing-wax*, *fat*, *indiarubber*, and *camphor*. Add to the first *methylated spirit*, to the second also *methylated spirit*, to the third *ether*, to the fourth *carbon-disulphide*, and to the fifth *chloroform*.

Observe that after a short time the shellac dissolves, after some hours the sealing-wax grows quite soft, and then mainly dissolves, leaving only a fine red powder (red lead, which is used to colour the red sealing-wax). The fat, indiarubber, and camphor dissolve in the respective liquids.

We learn from these experiments that—

Shellac	will dissolve	in	alcohol	(methylated	spirit).	
Sealing-wax	„	„	„	„	„	„
Fat	„	„	„	ether.		
Indiarubber	„	„	„	carbon disulphide.		
Camphor	„	„	„	chloroform.		

From these last experiments we may make the deduction that many bodies which cannot be dissolved by water may be taken into solution by other liquids.

Special names have been employed for certain solutions; thus, solutions prepared with spirit are known as **tinctures**, and strong solutions of sugar as **syrups**.

SOLUTION OF ONE LIQUID IN ANOTHER.—Many liquids will dissolve freely one in the other; thus, alcohol, glycerine, oil of vitriol, acetic and nitric acids will at once dissolve in water.

Experiment 43.—Place in a long glass tube a small quantity of coloured water, and then add about an equal bulk of methylated spirit, pouring the spirit carefully down the side of the tube.

Observe that the methylated spirit floats on the water, for it is lighter than water; but on shaking the contents they will completely mix, and do not separate again on standing.

We learn from this that these two liquids, alcohol and water, are soluble at once one in the other.

LIQUIDS INSOLUBLE IN ONE ANOTHER.—Some liquids do not behave with one another as the water and

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spirit in the last experiment, for they may be totally insoluble, and separate at once after even the most violent shaking, as is the case with mercury and water. Or they may remain mixed, but not dissolved, for varying lengths of time, as in the case of oil and water; the interval which elapses before they separate will depend upon the intimacy of the mixture and the fineness of the particles, For example, milk contains fat in suspension, not in solution, and it only floats to the surface as cream after the milk has been standing for some time, because the particles of oil are very minute and are closely entangled by the water particles.

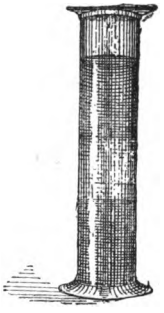

Fig. 18.—Jar containing mercury, water, and alcohol.

Experiment 44.—Pour into a long glass tube some olive oil, and then about an equal bulk of water, and well shake, then allow the mixture to stand; add about an equal volume of quicksilver and shake again.

Observe that the three liquids will not dissolve one into the other, but separate into three decided layers: the heavy mercury at the bottom of the tube, the water in the middle, and the oil floats on the top.

We learn from this that oil and mercury are both insoluble in water, and that one is lighter and the other heavier than water.

SOLUTION OF GASES IN WATER. — Many gases dissolve readily in water. The **spirit of hartshorn** is *ammonia gas* dissolved in water, and **spirits of salt** is really water holding in solution *hydrochloric acid gas.* **Soda water** and other effervescent drinks owe their effervescence to the presence of *carbonic acid gas in solution.* All natural waters contain gases derived from the

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atmosphere in solution, and it is from this dissolved air that fishes are able to obtain the oxygen which they require as the water passes through their gills.

Experiment 45.—Completely fill a flask with *cold tap water*, as shown in Fig. 19, provide it with a well-fitting cork and delivery tube, and see that the cork is in direct contact with the water. Allow the delivery tube to dip into a glass trough, and place the open end beneath an inverted test-tube, which has been filled with water as shown. Heat the flask.

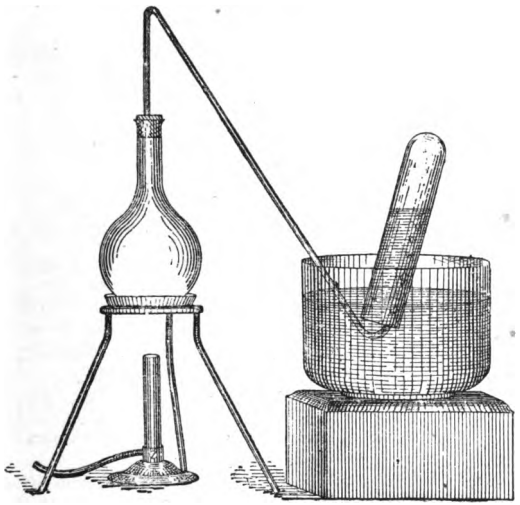

Fig. 19.—Apparatus for showing that water contains air in solution.

Observe that long before the water boils bubbles of an invisible gas rise through the test-tube and displace the water. The gas on examination proves to be air.

We learn from this that ordinary water contains air dissolved in it, and that this air may be expelled by heat.

Other experiments in later lessons will demonstrate the facts that other liquids than water will dissolve gases, and that, as a rule, **cold liquids can dissolve more of any gas than hot liquids.**

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SUMMARY TO CHAPTER IV.

SOLUTION - - -
- *Solvent*, a dissolving agent.
- *Soluble*, the body which is dissolved.
- *Insoluble*, a substance which will not dissolve.
- Quantity of matter dissolved depends upon the temperature.

EVAPORATION -
- Method of recovering solids from solutions by driving off the liquid by heat.

FILTRATION - -
- *Filtrate*, the liquid which passes through the filter.
- *Residue*, that which remains on the filter.
- *Soluble bodies* pass through a simple filter, and *insoluble substances* remain in the filter.

DECANTATION -
- May be resorted to where the *insoluble* body consists of *heavy particles*.

CRYSTALLIZATION - - - -
- A method of regaining dissolved solids, in the form of *crystals*, or preparing crystalline forms by *sublimation*.

NATURAL WATERS - -
- *Hard and soft:* The former contains salts in solution which curdle soap, the latter will not curdle soap. *Temporarily hard* water may be softened by boiling. *Permanent hardness* cannot be removed by boiling.
- *Sea water:* Saline and hard, rich in soluble salts.
- *River water and lake water:* More or less turbid through matter held in suspension; usually contain matters in solution; hard or soft according to the nature of the rocks of the district. *Thames water* hard, *Loch Katrine water* soft.
- *Well water:* That from shallow wells often contains much organic matter in solution. Deeper well water is clear and, as a rule, well fitted for drinking purposes.
- *Spring waters:* Usually hard and charged with saline substances in solution. When the salts present impart decided flavour, the water is known as *mineral water*.

SOLVENTS - - -
- Water, the general solvent; but other liquids may be employed, such as spirit, ether, chloroform, and turpentine. Some liquids will dissolve in others, alcohol in water; others are quite insoluble, oil, mercury, and water.
- Most gases will dissolve in water, and some in other liquids; some gases dissolve very readily in water, for example, carbonic acid gas, ammonia, and hydrochloric acid gas.

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QUESTIONS ON CHAPTER IV.

1. Explain the following terms:—Solvent, soluble, insoluble, simple solution, chemical solution, filtration, decantation.

2. How would you separate—(1) sand from sugar; (2) chalk from common salt; (3) charcoal from nitre; (4) sulphur from water?

3. What do you understand by the term evaporation, as employed by chemists?

4. How would you obtain the crystals of any body from a solution of the substance?

5. How would you separate chalk from sugar, and soda from lampblack? Sketch the whole of the apparatus you would employ.

6. What is the hardness of water due to?

7. How are natural waters classified?

8. Point out the chief differences between sea water and river water.

9. How does rain water become impure?

10. Of the natural drinking waters, which are the best for drinking purposes?

11. Explain the difference between water which is permanently hard, and that which is temporarily hard. What is the cause of the difference?

12. How would you ascertain whether a sample of water contained any matter in solution?

13. Describe the apparatus you would employ to make dirty water clear.

14. Explain how you would demonstrate whether a sample of tap-water contained any matter in solution.

15. Write out the names of *ten* common soluble bodies and five common insoluble substances.

16. What are mineral waters, and how are they classified?

17. How may dirty water be made clear?

18. Name some of the important salts found in sea water.

19. How would you prove that ordinary tap water contains air in solution? Mention any other gases which are soluble in water?

20. Name some liquids which are soluble in each other.

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Preparation for Fifth Lesson.

APPARATUS AND REAGENTS REQUIRED.

Bell jars, pneumatic trough of water, popgun or syringe, flask and safety funnel, bulb tube, syphon tube as in Fig. 24, mercury, flask, balance, tumbler of water, barometer tube or glass tube about 33 inches in length closed at one end, lime-water, litmus solution, candle and two lamp glasses, phosphorus, mercuric oxide, retort, potassic chlorate, manganese dioxide, gasholder, solid pyrogallic acid, sodic hydrate, piece of carbon, sulphur, iron wire, zinc foil, six gas jars and plates, deflagrating spoons, cedar splints, dry calcium chloride, and dinner plate.

EXPERIMENTS TO BE PERFORMED.

"That air occupies space may be shown by plunging a bell jar into a vessel of water.

Close the short limb of a syphon tube containing air, and compress the air in the short limb by pouring in mercury.

Weigh a flask full of air, and fitted with a cork, caoutchouc tube, and clip; exhaust by sucking out the air, and weigh again.

Show that a lighted candle is soon extinguished when burnt under a glass resting on the table, but that it continues to burn if air be supplied by supporting the lamp glass on two pieces of wood.

Burn phosphorus in a bell jar over water, and show the diminution of air. Ignite phosphorus, place it in the remaining gas.

Burn some phosphorus under a dry bell jar to show the compound of phosphorus and oxygen which is formed.

Fix a pellet of phosphorus on the end of a piece of wire, and place in a test-tube graduated with a slip of paper and inverted in a tumbler of water to show that at ordinary temperatures it combines with the oxygen of the air and removes it, so that by measuring the volume of gas left the amount of oxygen contained in air can be roughly determined.

Burn charcoal and sulphur in oxygen, and call attention to their disappearance. Demonstrate by lime-water and by litmus paper respectively that a new body is in each case formed.

Burn iron powder (*Ferrum reductum*) on a scale pan of a balance to show that an increase of weight occurs.

A glass or metal vessel filled with ice or cold water can be used to show the condensation of moisture upon it.

Place calcium chloride on the pan of a balance to show the gradual increase of weight which occurs by absorption of aqueous vapour from the air."—*Science Department's Syllabus.*

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CHAPTER V.

OXYGEN AND NITROGEN.

Air surrounds the Globe—Wind is Air in Motion—Air occupies Space—The Bulk of any Quantity of Air is much changed by Temperature and by Pressure—Air has Weight—The Necessity of Air for Animals and Plants—Bodies when burning require Air—Air a Mixture of two Gases—viz., Oxygen and Nitrogen—The Proportion of Nitrogen to Oxygen—Oxygen the Active Body in Air—Bodies burn in Oxygen, and more Brilliantly than in Air—The Combination of Oxygen with Iron and with other Bodies—Increase of Weight of Bodies which unite with Oxygen—Nitrogen does not combine directly with Bodies—The nearly constant Composition of Pure Air—Presence of other Gases in small amount in Air—Water in the Air as a Gas—The drying up of Water.

THE name **AIR** is given to the atmosphere which completely **surrounds the globe.** It is a mixture of invisible gases which is without odour or taste, but still we may know that it exists and is matter; for although we cannot see it, yet we feel its pressure when the wind blows and its resistance when we move quickly through it, and we may prove by very simple experiments that it **occupies space** and **possesses weight** like all other forms of matter.

Experiment 46.—Plunge a glass bell jar or large glass into a vessel of water.

Observe that the water rises only very slightly in the bell jar, and that as the glass is pressed deeper into the water more force has to be employed; when the hand is removed the glass springs up in the water.

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We learn from this that air occupies space to the exclusion of the water, and that it is necessary to employ force in order to press the air into a smaller space; and we also see that the air must be **elastic,** for the glass springs up from the water when the pressure is removed.

A body is said to be **elastic** when it resumes its original form after it has been made by any means to undergo a change in form. The change may be brought about by stretching, twisting, bending, compression, etc. In our last experiment we reduced the volume, or space occupied by the air, by compression; but as soon as the pressure was reduced or removed, the air regained its former volume and pressed the water out of the glass.

Fig. 20.—Air occupies space and resists pressure.

A vessel is usually spoken of as "empty" when it contains only air, but that it is not really empty but contains matter in a gaseous form we have proved in our experiment, for since it offered resistance to the entrance of the water it must have been full of something, and that which filled it we call air.

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Experiment 47.—Close the end of a pop-gun, or the nozzle of an ordinary syringe, or one end of a piece of glass tubing, and then compress the air contained within by means of the piston.

Fig. 21.—Apparatus to illustrate elasticity of air.

Observe, if the piston fits well, that the air is compressed, and when released the piston flies back. On the other hand, if the piston is drawn out, so that the air is allowed to take up a larger space on releasing the piston, it returns to its former position.

We learn from this that air is elastic, for it tends to return to its original volume when once the form has been changed.

Air when heated increases in volume. This expansion of air under the influence of heat may be illustrated by a very simple experiment.

Experiment 48.—Fit up a small *flask*, containing *air*, with a cork provided with a safety funnel, as shown in Fig. 22, pour into the safety tube a small quantity of coloured water. Now heat the flask.

Observe that when the air is heated it expands and raises the coloured liquid in the tube, and that on removing the source of heat the liquid falls again.

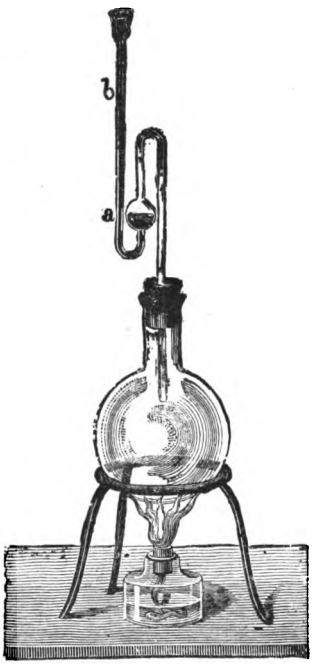

Fig. 22.—Expansion of air.

We learn from this that air when heated expands, and when cooled contracts. Its bulk, therefore, depends upon its temperature.

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That a very **slight change in temperature** is sufficient to bring about a change in bulk of air may be shown by means of the glass bulb-tube, shown in Fig. 23. This tube has a fine bore, and is provided with a bulb; a drop of coloured water is introduced as shown at *a*. When the bulb is held in the warm hand, the drop of liquid is pressed upwards by the expanding air, indicating that the enclosed gas undergoes a change in volume, even when the change of temperature is very slight.

Fig. 23.—Bulk of air increased by very slightly heating it.

Experiment 49.—Fill a flask with coal-gas, and fit it up as in Fig. 22, and heat it.

Observe that the expanded coal-gas escapes and is collected in the test-tube; when a burning match is brought to the mouth of the test-tube the gas takes fire.

We learn from this that coal gas expands when heated. Many other experiments of a similar nature have led to the conclusion that **the volume of a gas varies with its temperature.**

We saw in experiment 46 that the water rose slightly as the glass was pressed downwards, and we learnt that greater force was required to press the air in the glass into a smaller space to make room for the rising water. Compare this fact with the results obtained in the next experiment.

Experiment 50.—Fit up a syphon tube as shown in Fig. 24, and pour a small quantity of mercury down the long limb, so that a certain bulk of air is made a prisoner in the short closed limb of the tube between 0 and 25. Now pour more mercury into the long limb.

Observe that as the height of the mercury in the long limb increases, the liquid rises in the short limb, pressing the confined air into a smaller

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space. When the difference in level between the mercury in the two limbs is about thirty inches, the air will occupy half its original volume. Now pour out some of the mercury and observe that the volume of gas increases.

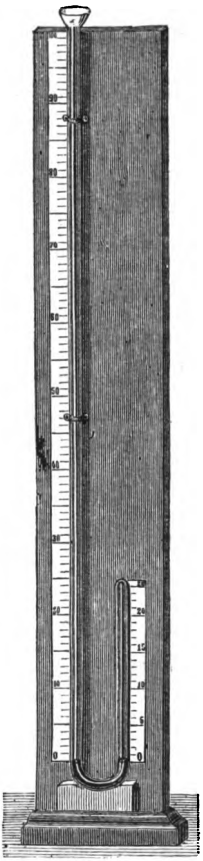

Fig. 24.—Apparatus for showing that the volume of air varies with pressure.

We learn from this that the **volume of air depends upon the pressure** to which it is subjected, for its bulk decreases with increased pressure and increases with decreased pressure.

Experiment 51.—Provide a large light flask with a sound cork, and carefully weigh it; then drive out as much of the air as possible by heating the flask, or suck out the air; then carefully weigh again.

Observe that the flask decreases in weight; this loss of weight represents the air which has been removed.

We learn from this that **air has weight,** although we may be inclined to lose sight of the fact because we are so little inconvenienced by it.

One pound of air occupies about 13 cubic feet, although, as we have seen in experiments 49 and 50, the volume will vary with pressure and temperature. At ordinary atmospheric pressure and temperature, 100 cubic inches of air weigh 31 grains.

Since air has weight it must exert pressure, and therefore all things on the earth have to bear this atmospheric pressure, which amounts to about 15 lbs. on every square inch at the sea-level.

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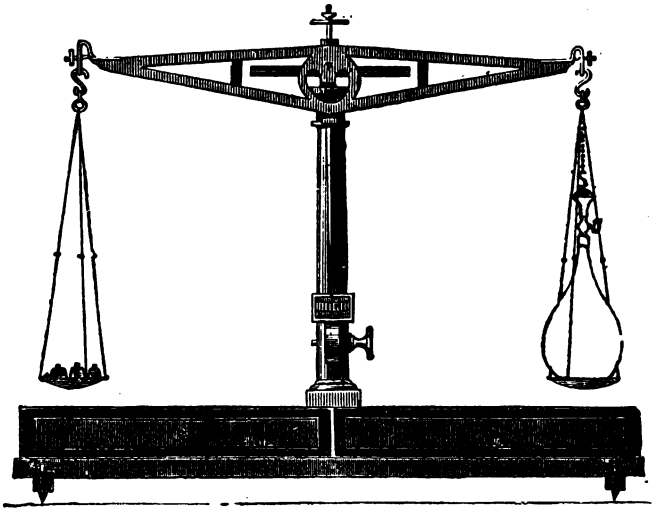

Fig. 25.—Apparatus for weighing air.

Experiment 52.—Completely fill an ordinary tumbler with water, and carefully cover it with a piece of ote paper. See that the paper touches the water and the rim of the glass. Place one hand flat on the paper, and invert the glass as shown in Fig. 26.

Observe that the paper remains as if glued to the glass and the water in the glass when the vessel is inverted, and as long as air is not admitted.

We learn from this that air exerts pressure.

Experiment 53.—Fill a glass tube, closed at one end and about 32 inches in length, with mercury, and close the end with the thumb and invert it in a cup of mercury.

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Observe that on removing the thumb the mercury falls until the column stands at about 29½ inches above the level of the liquid in the dish

We learn from this that the pressure of the air is sufficient to support a column of mercury about 30 inches in height.

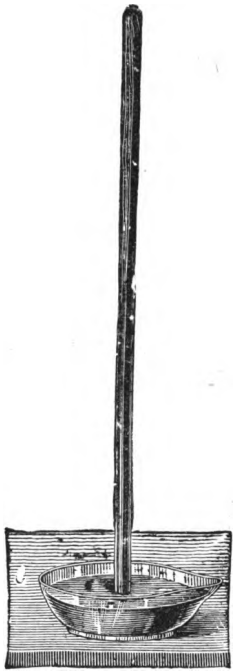

Fig. 27.—Barometer tube to illustrate pressure of the atmosphere.

Necessity of Air for Animals.—All living animals with lungs constantly introduce fresh air into those organs during the process of respiration, and just as constantly cast out vitiated air. That air is absolutely necessary for the life of an animal has been shown by placing a bird under the receiver of an air pump, and drawing out the air. In such an experiment, a small bird has been killed at once.

Some of the oxygen of the air breathed in is absorbed by the blood in the lungs. At the same time the carbonic acid gas is increased more than one hundred times, for whilst pure air only contains ·04 per cent. of carbonic acid gas, expired air contains about 4·5 per cent. of this gas. Expired air always contains some organic matter, and it is saturated with water vapour.

The presence of carbonic acid gas in expired air may be demonstrated by breathing into some clear lime-water

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contained in a glass, as in Fig. 29, when it will be observed that the lime-water becomes milky.

That moisture is largely present in expired air, may be proved by breathing on any cold polished surface, when the brightness will be dimmed by the water vapour which is condensed.

It is the presence of this expired carbonic acid gas, associated with organic matter, which makes the air of overcrowded rooms, workshops, and dwellings so harmful to the inmates. By **ventilation** we aim at promoting the rapid removal of these impurities which act so perniciously on the health.

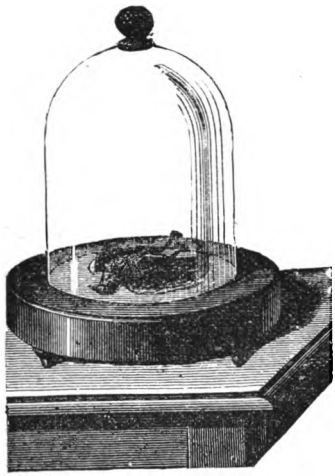

Fig. 28.—Air necessary for life.

Experiment 54.—(*a*) Fit up a wash-bottle, place a small quantity of lime-water in the bottom of flask, and see that the long tube dips well into the lime-water.

(*b*) Place the short tube in the mouth and draw air from the bottle so that the gas taken into the lungs must pass into the flask through the lime-water. Do this for a minute.

(*c*) Then apply the mouth to the long tube, and blow the expired air through the lime-water for the same length of time.

Observe that in the first case the lime-water remains clear, and in the second it becomes milky at once.

We learn from this that the air breathed in must differ from the air expired, as far at least as the proportion of carbonic acid gas is concerned.

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All animals absorb oxygen: **air breathers** take it directly from the air; those animals which breathe by gills obtain it indirectly from the air, for they obtain it from the air which is dissolved in water. Hence it becomes necessary to supply fishes kept in an aquarium with fresh water frequently, or to drive air into the water constantly.

Necessity of Air for Plants. — Plants of all kinds absorb from the air carbonic acid gas as well as oxygen. Under the influence of light, and by the action of the green colouring matter (chlorophyll) which they contain, they decompose the carbonic acid gas and liberate oxygen.

Fig. 29.—Lime-water experiment.

Experiment 55.—Place a few sprigs of some fresh green plant, suc as watercress, in a glass jar provided with a stopper, as in Fig. 30. Fill the jar with water, cover it with a glass plate, and invert it in a vessel of water, then remove the plate. Expose the jar to the bright daylight.

Observe that small bubbles of gas appear on the leaves of the plant; they increase in size, and finally are liberated and rise through the water and collect in the upper part of the jar. When this gas is examined by means of a glowing splint, it will be found to answer the tests for oxygen gas.

We learn from this experiment that green plants, under the influence of light, liberate oxygen.

Bodies when burning require air to promote the chemical changes which take place during combustion, as

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we saw in experiments 8 and 9, and to carry away the products of combustion.

Experiment 56.—Place a lighted candle under a bell jar.

Observe that the candle flame is soon extinguished.

We learn from this that a candle will not continue to burn in a limited supply of air.

Experiment 57.—Place a piece of lighted candle upon the bench, and put over it a bell jar from which the stopper has been removed, and let the jar rest upon two pieces of wood, so that air can enter below and the products of combustion escape above.

Fig. 30.—Candle flame extinguished when the supply of air is cut off.

Observe that the candle in this case continues to burn until the whole is used up.

We learn from this that a continuous supply of fresh air is necessary when bodies are burning.

Experiment 58.—Take a piece of phosphorus about the size of a small pea, and place it in a crucible floating in a vessel of water; ignite the phosphorus and cover it with a glass bell jar.

Observe that as the phosphorus burns the water rises in the jar, and that the vessel becomes filled with white fumes. After standing for a short time the fumes disappear, and the jar is seen to contain an invisible gas. If a piece of burning phosphorus be placed in this remaining gas *the flame is at once extinguished.*

We learn from this that phosphorus in burning uses up part of the air, for the water rises, showing that the gaseous matter has decreased in the jar. Since the phosphorus will not burn in the remaining gas it cannot be air, and the air in making the original phosphorus burn gave

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up that part of its matter which supports combustion. The gas which is left in the jar, in which the phosphorus will not burn, is termed **nitrogen,** and that which has been used up is **oxygen.** With this oxygen the phosphorus has combined chemically, forming the compound body known as *oxide of phosphorus;* the fumes which filled the jar at first, and which were absorbed by the water consist of clouds of the finely divided phosphorus oxide.

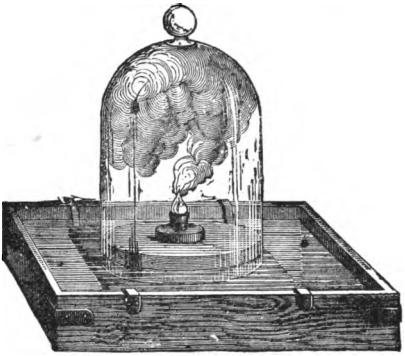

Fig. 31.—Phosphorus burning in air under a bell glass over water.

Thus then we see that **air is a mixture of two gases, namely, oxygen and nitrogen.**

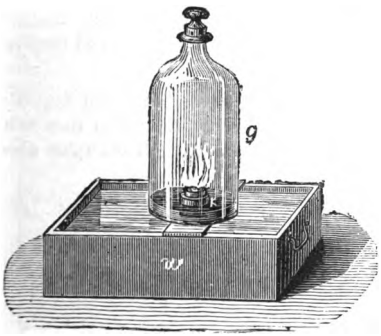

Fig. 32.—Proportion of nitrogen to oxygen in air.

Experiment 59. — Place a piece of phosphorus on a dry tile, and touch it with a hot wire; then quickly invert over the burning phosphorus a cold, dry bell jar; raise the jar a little on one side from time to time to admit more air.

Observe that a white snow-like powder is produced, which is deposited upon the tile and in flakes upon the glass. A little of it cast into water dissolves at once with a hissing noise.

We learn from this that oxide of phosphorus is a white solid body, very soluble in water.

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The proportion of nitrogen to oxygen in air by volume may be demonstrated roughly by the two following experiments :—

Experiment 60.—Fix a pellet of phosphorus on the end of a piece of wire, and place in it a test-tube graduated with a slip of paper into five parts, and inverted in a vessel of water.

Observe that, at the ordinary temperature of the air, in about half an hour the water will rise through one division of the test-tube.

We learn from this that phosphorus combines with the oxygen of the air at ordinary temperatures, and that about *four-fifths of that air consists of nitrogen and one-fifth oxygen by bulk.*

Experiment 61.—(*a*) A glass tube about 30 inches in length should be sealed at one end, and the other closed with a sound, well-greased cork.

(*b*) The tube should then be divided into five parts by means of small indiarubber rings cut from a piece of wide tubing.

(*c*) A strong solution of pyrogallic acid should then be prepared, and a strong solution of caustic potash.

(*d*) Quickly introduce about 10 cc. of the pyrogallic acid solution into the tube and add a few drops of caustic potash. Replace the cork and shake well.

(*e*) After shaking the pyrogallic acid with the air for a few minutes, place the mouth of the tube well in a vessel containing water, and carefully withdraw the cork under water.

Observe that the pyrogallic acid solution turns brown on agitating with the air. This is due to absorption of oxygen. Note also that when the mouth of the tube is opened, the water rises until it occupies about one-fifth of the tube.

We learn from this that the oxygen, which has been removed by the pyrogallic acid, and the place of which has now been taken by water, must have made up one-fifth of the total quantity of air in the tube.

OXYGEN is the active body in air, and originally it was called *vital air*, because the chemists who discovered it found that no animal could live without it. Oxygen is

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the most abundant element in the earth; it enters largely into the composition of all animal and vegetable substances; it forms eight-ninths of water by weight, and, as we have just seen, forms one-fifth of the bulk of the atmosphere. It exists free, but mixed with nitrogen in air; in rocks, water, and living things it is found in combination with other elements. Since this gas is so widely

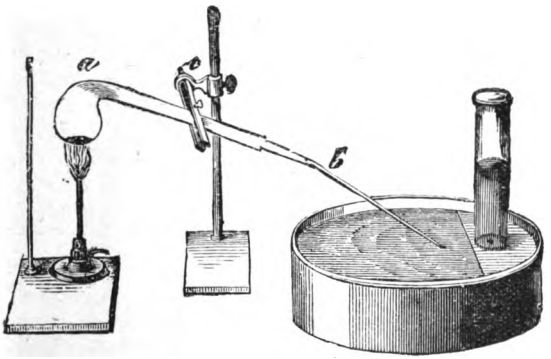

Fig. 33.—Preparation of oxygen from mercuric oxide.

distributed, and so abundant in nature, it would be interesting to study some of its more important properties. It cannot be prepared directly from air, for the nitrogen with which it is mixed is such an inert body that it is exceedingly difficult to remove it; but indirectly it may be prepared from air by first fixing the oxygen by means of mercury. When this metal is heated nearly to boiling for some hours in contact with air, it becomes slowly covered with a reddish film of oxide, and increases in weight. When this mercuric oxide is heated to a higher temperature it gives up oxygen.

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Experiment 62.—Place a little mercuric oxide in a hard glass retort, as shown in Fig. 33, fit to the neck of the retort a delivery tube, so that it dips under water in a pneumatic trough. Invert a jar full of water over the mouth of the delivery tube, as shown in the sketch, and strongly heat the retort.

Observe that bubbles of gas come off which displace the water in the gas tube, and the neck of the retort becomes coated with globules of metallic mercury at *a*. The collected gas will not burn in air, but rekindles a glowing cedar splint.

We learn from this that oxygen may be prepared

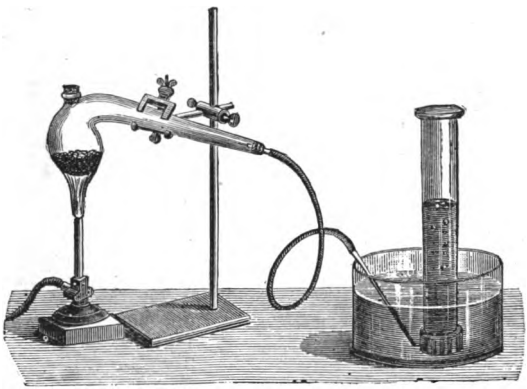

Fig 34.—Preparation of oxygen from *potassium chlorate* and *manganese dioxide.*

from mercuric oxide, and it is a powerful supporter of combustion.

Where large quantities of oxygen are required it is much more convenient to prepare it from potassic chlorate, which is very rich in this element, and which parts with the whole of it when it is strongly heated.

Experiment 63.—Prepare a mixture of five parts by weight of *potassic chlorate* and two parts by weight of *manganese dioxide.* Place the mixture in a retort, flask, or large test-tube, fitted with a sound cork and wide delivery tube.

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Fix the retort on a stand, and let the delivery tube dip under water in a pneumatic trough, as shown in Fig. 34.

Heat the retort, and collect six or seven jars of the gas which escapes.

Observe that the mixture, on heating, does not fuse and that an invisible gas displaces the water in the gas tubes.

We learn from this that oxygen may be prepared with ease, and in large quantities, from a mixture of manganese dioxide and potassic chlorate.

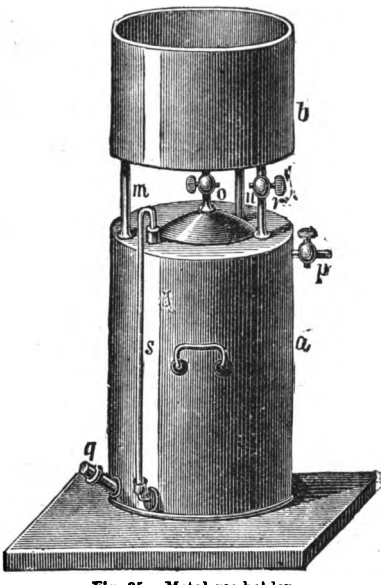

Fig. 35.—Metal gas-holder.

Oxygen obtained by the above method is best stored for use in a gasholder constructed of metal, as in Fig. 35. To charge the gasholder with gas we proceed as follows:—Fill the large vessel *a*, Fig. 35, with water by pouring water into *b* and turning on the stopcock *r*, which is fixed on a tube which runs quite to the bottom of *a*; the water then runs from *b* to *a* if the stopcock *p* has been previously turned on so that the air may escape. The graduated glass tube *s* informs us of the level of the liquid in the gasholder, and therefore of the amount of air present. Water should be passed from *b* to *a* until the tube *s* indicates that *a* is quite full of water. The stopcocks should now all be turned off, and the screw-cap at *q* removed. The pressure of the air prevents the liquid running out. A glass tube from an oxygen generating apparatus should next be passed into *a*, through the opening at *q*, and oxygen will rise through the water to drive out the liquid. When the holder *a* is quite full the screw-cap should be replaced at *q*. The gas may be used from the holder directly by pouring water into *b* and turning on the stopcocks *r* and *p*. The water passes down from *b* through *r* into *a*, and drives the oxygen out through *p*. A gas jar may be filled by placing it over the central tube *o*, which opens into *b* and just passes into the top of *a*, below. The cock *p* should be turned off, *b* should be filled

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with water, a gas jar placed over *o*, then *r* partially turned on, water then runs down the right-hand tube and drives out oxygen through the open stopcock *o* into the gas jar which has been placed in *b*. It will be observed that in this case *b* serves as a pneumatic trough. It is supported by the upright rods marked *m n*, and partially by the tube *r*.

Experiment 64.—Introduce into a jar of oxygen a smouldering splint or piece of brown paper.

Observe that the gas is invisible and inodorous, and notice that the gas supports combustion, for the chemical action becomes so rapid, and the amount of heat produced so great, that the smouldering wood bursts forth into a bright flame.

We learn from this that oxygen is a more powerful supporter of combustion than air, for the body only smoulders in air, but burns brightly in oxygen.

Experiment 65.—Introduce into a second jar of oxygen a little clear *lime-water*, taking care not to remove the covering plate more than is absolutely necessary. Hold the plate firmly on the jar and well shake.

Observe that the lime-water remains quite clear, thus proving that oxygen has no action on lime-water. Now burn a taper in the jar and again shake up the lime-water with the remaining gas, and notice that it turns quite milky.

We learn from this that some oxygen disappears when wood is burnt in it, and produces the gas which turns lime-water milky, which is known as **carbonic acid gas.**

Experiment 66.—Introduce into a third jar of oxygen a small piece of carbon which has been placed in a deflagrating spoon. The carbon must be heated in a gas flame until it glows; in this condition it should be plunged into the jar of oxygen.

Observe that the carbon burns brightly, giving out a brilliant light. If a little lime-water be placed in the jar of oxygen before the experiment, it remains quite clear, but after the combustion of the carbon, it slowly turns milky.

We learn from this that oxygen is a powerful supporter of combustion, and that when carbon is burnt in oxygen **carbonic acid gas** is produced.

Experiment 67.—Pour a little *blue litmus solution* into the fourth jar of oxygen. Then place a small piece of sulphur in a deflagrating spoon, ignite it, and introduce it into the jar; after the flame of the sulphur has gone out, remove the spoon, cover the jar with a plate, and place it on one side.

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Observe that the oxygen has no action upon blue litmus, that the sulphur burns with a bright flame, and that the jar becomes filled with a heavy strong-smelling gas, which turns the blue litmus red.

Fig. 36.—Sulphur burning in oxygen.

We learn from this that sulphur burns in oxygen, giving rise to a new gaseous substance called **sulphur dioxide.**

The last three experiments have demonstrated that the non-metallic bodies, of which we have types in carbon and sulphur, which will burn in air, undergo more brilliant combustion in oxygen. If in each case the resulting bodies are examined by means of moistened *blue litmus paper*, we shall find that the substances produced possess acid properties; in the case of carbon dioxide the blue litmus is only slightly affected, being turned claret colour, but in the other case it is at once turned decidedly red.

Fig. 37.—Iron wire burning in oxygen.

We shall next demonstrate by a few experiments that oxygen will also combine very readily with common metals.

Experiment 68.—Warm a strip of zinc foil, and bring the tip in contact with some sulphur; ignite the sulphur and quickly introduce the zinc into the fifth jar of oxygen.

Observe that the burning sulphur raises the temperature of the zinc, so that it ignites and burns with a dazzling white light. The burning zinc yields a powder which is yellow whilst hot and white when cold.

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We learn from this that metallic zinc will burn in oxygen to form **zincic oxide.**

Experiment 69.—Wind a thin piece of iron wire round a pencil so as to form a spiral, fasten one end to the plate of a deflagrating spoon, and to the other end attach a small piece of carbon; ignite the carbon and plunge it into the sixth jar of oxygen.

Observe that after the carbon, which serves to make the iron hot, has burnt, the iron itself takes fire and burns with brilliant scintillations, and the sparks which fly off should be allowed to fall into a little water placed at the bottom of the jar.

We learn from this that the metal iron will combine with oxygen to form solid **ferric oxide.**

Experiment 70.—Place about half an ounce of iron powder (*Ferrum reductum*) in a watch glass and weigh, then ignite the powder with a red-hot wire and carefully weigh again.

Observe that the body increases in weight on burning.

We learn from this that iron in uniting with oxygen during combustion has increased in weight. It is true in all cases that the *products resulting from the combustion of bodies weigh more than the bodies burnt.*

Nitrogen we prepared from air in experiment 57. This body is an invisible gas, which exhibits little tendency to combine with other substances. It will neither burn nor support combustion; and as animals cannot live in it, the name **azote** (from *a*, not, and *zoe*, life) has been given to it. Mixed with oxygen in air, it serves to reduce the activity of that body, rendering the air less stimulating to animals and moderating oxidation and combustion.

The composition of pure air is nearly constant, in spite of the fact that the two chief component gases are mixed and not united chemically; but it is well to remember that other gaseous bodies are always present with the oxygen and nitrogen.

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Experiment 71.—Place about half an ounce of dry powdered calcium chloride in a watch glass, and counterpoise it with small shot on the pan of a balance.

Observe that after a time the salt becomes moist in appearance, and increases in weight by absorption of aqueous vapour from the air.

We learn from this that water exists in the air in the gaseous form.

The deposition of dew and formation of rain and snow are proofs of the presence of invisible water in the air. The drops of water or hoar-frost deposited upon the outside of a metal or glass vessel containing a freezing mixture, also demonstrate the presence of moisture in the air. And the **drying up of water** proves that it must pass into the air in an invisible form.

Experiment 72.—Pour a little clear *lime-water* into a shallow dinner plate, and allow it to remain exposed to the air for some time.

Observe that a whitish film of carbonate of lime is formed upon the surface.

We learn from this that carbonic acid gas exists in air.

In addition to aqueous vapour and carbonic acid gas and ammonia, other gases exist in air; but when these bodies have been removed, the following represents the result obtained after many experiments:—

Oxygen, by vol. 20·9, by weight, 23·1; nitrogen, by vol. 79·1, by weight 76·9.

The following shows the proportion of gases in one hundred parts of pure air:—

Nitrogen, 77·95; oxygen, 20·61; water vapour, 1·40; carbonic acid gas, ·04; ammonia, ozone, oxides of nitrogen, other bodies, traces.

In the air of towns and manufacturing districts, such bodies as sulphur dioxide and sulphuretted hydrogen have been detected; and associated with these in the air of densely-populated districts we always find considerable quantities of carbon, organic and inorganic dust.

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SUMMARY TO CHAPTER V.

AIR.

- **Physical Properties** - - *Occupies space, possesses weight*, offers *resistance* to pressure; *elastic*. Volume varies with *temperature*, expands when heated, contracts when cooled. Volume varies with *pressure*, *density* increased by increased pressure.
- **Action of Animals and Plants on Air** — Animals take oxygen from air and give off into the air carbonic acid gas. Plants tend to reduce the carbonic acid gas in air under the influence of sunlight.
- **Air necessary for Combustion** — Bodies cannot burn without a constant supply of air.
- **Composition** - — Consists principally of oxygen and nitrogen, with small quantity of carbon dioxide, etc.
- **Properties** - — Will not burn, but is the great supporter of combustion; and oxidation, due entirely to the oxygen present, the nitrogen of the air taking no part in this process; is a mechanical mixture.
- **Analysis** - - — **Volumetric** shows it consists of 20·9 per cent. of oxygen and 79·1 per cent. of nitrogen. **Gravimetric** proves that the oxygen in air is mixed with nitrogen in the proportion of 23·1 parts of the former to 76·9 parts of the latter by weight.

OXYGEN.

- **Occurrence in Nature** - - — Makes up one-third to one-half of the weight of the *earth's crust*, forms eight-ninths of weight of *water*, and makes up one-fifth of the bulk of the *atmosphere*. It is a constituent of all *plants* and *animals*, and exists in nearly all *minerals*.
- **Properties** - — Colourless, tasteless, inodorous gas; powerful supporter of combustion, but does not burn in air. Combines with bodies to form *oxides*, and this leads to increase of weight of a body during combustion.

NITROGEN.

- **Occurrence in Nature** - - — Makes up four-fifths of air by volume, and is present in all living things.
- **Properties** - — Colourless, tasteless, inodorous gas, will not burn, and non-supporter of combustion.

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QUESTIONS ON CHAPTER V.

1. How would you prove that air has weight?

2. Show how you would demonstrate the more important physical properties of air.

3. What are the effects produced upon air (*a*) by a change of temperature, (*b*) by a change in pressure?

4. How would you prepare oxygen?

5. What are the resulting products when carbon, sulphur, and iron are separately burnt in oxygen gas?

6. How would you prove that bodies increase in weight when burnt?

7. State the chief properties of nitrogen, and explain how this body usually occurs in nature.

8. Why is the flame of a taper extinguished in nitrogen gas, and why does it continue to burn in air?

9. Describe three experiments, the best you can think of,—one to prove that the air contains oxygen, another that it contains carbon dioxide, and the third that it contains aqueous vapour.

10. How would you prove that in expired air there is carbon dioxide and aqueous vapour?

11. Describe the chemical changes produced on the atmosphere by living animals and plants, and mention how these changes may be demonstrated by experiment.

12. Describe two experiments and sketch the necessary apparatus for obtaining from the air (*a*) nitrogen, (*b*) oxygen.

13. By what experiments can you prove that the air contains four-fifths of its volume of nitrogen?

14. What other bodies are sometimes met with in air besides oxygen and nitrogen?

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Preparation for Sixth Lesson.

APPARATUS AND REAGENTS REQUIRED.

Drinking water, distilled water, ice, flasks, glass tubes, corks, balance and weight, thermometers, sawdust, beakers, kettle and spirit lamp, copper sulphate, retort, alcohol, carbonate of soda, combustion tubing, and apparatus as shown in Fig. 46.

EXPERIMENTS TO BE PERFORMED.

"Illustrate the characteristic properties of ice, water, and steam.

Show that equal volumes of hot and cold water do not counterbalance one another.

Fill a flask to the bottom of the neck with cold water, and then heat to slow expansion of the water.

Show current by heating a large flask of water.

To illustrate distillation distil water containing copper sulphate. Show Liebig's or other forms of condensers.

Show the mode of determining the boiling point of a liquid.

Show that the temperature remains constant, and that on dissolving substances in water the boiling is raised."—*Science Departments' Syllabus.*

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CHAPTER VI.

PHYSICAL PROPERTIES OF WATER.

Water: Its three States—Expansion of Water by Heat—Equal Volumes at Different Temperatures have not the same Weight—Formation of Currents in Water by Heating—Boiling Point—Increase of Volume on Conversion of Water into Steam—Distillation—Pure Water.

Water covers the greater portion of our globe. It forms part of all animals and vegetables. It is of so much use to man for personal, domestic, and commercial purposes, and it is such a valuable agent to chemists, that the importance of a knowledge of its **physical properties** and **chemical composition** at once becomes apparent. At the ordinary temperatures in this country pure water is a clear, transparent liquid, without taste or smell. It is a most important solvent agent, and in consequence of the high degree in which it possesses this property, it is never met with in nature in a state of purity; for no sooner is it formed in the clouds than it dissolves atmospheric gases, and thus during its descent it gains various substances from the air, and on reaching the earth continues its solving action on the rocks with which it comes in a retort beaker of mixture of Water is met with in the three well-known **states**—solid as **ice**, liquid as **water**, vaporous

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When cooled from the ordinary temperature of the air down to 32° F. or 0° C., it passes into the solid condition, in which form it is of course known as ice. In this condition it is bulk for bulk lighter than water; **ice therefore floats on water.** Most bodies contract and become heavier volume for volume as their temperatures fall.

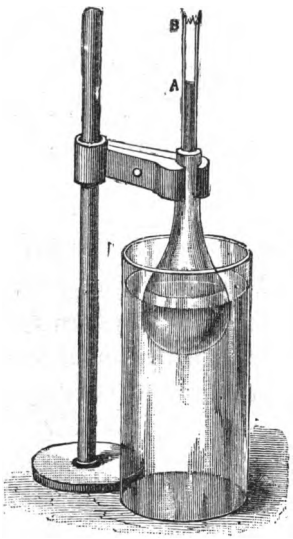

Fig. 38.—Expansion of heated water.

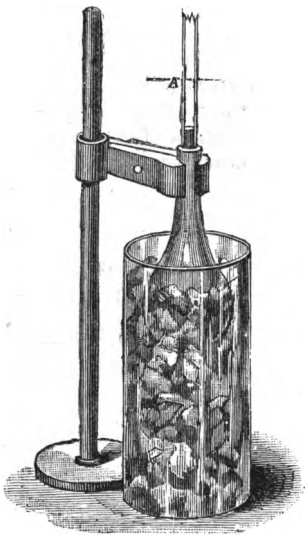

Fig. 39.—Contraction of cooled water.

If in this respect water obeyed the general rule, then, on assuming the solid condition, it would necessarily become heavier bulk for bulk and sink in water. Water cooled [illegible] 100° C. to 4° C. contracts, thus following the rule; 4° C. it reaches its **maximum density,** and it [illegible]rom this temperature down to 0° C., at which point [illegible]d.

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Experiment 73.—Fit up a flask with a cork and long narrow tube, as shown in Fig. 38. Fill the flask with cold water, and mark the point on the tube to which the liquid reaches. Now plunge the flask into a beaker of hot water.

Observe that as the vessel becomes heated the water rises in the tube.

We learn from this that **heat causes water to expand,** and since the same weight of water now occupies a larger space, it follows that the hot water must be lighter bulk for bulk than the cold water.

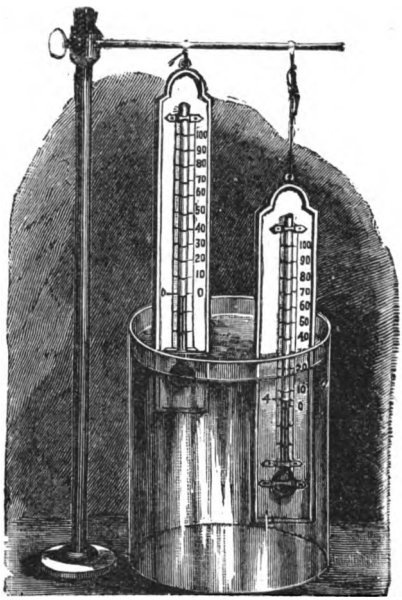

Fig. 40.—Apparatus to illustrate maximum density of water is reached at 4° C.

Experiment 74.—Now place the flask in a beaker containing some snow or broken ice.

Observe that the level of the liquid falls until it reaches the marked point, and that as it grows colder it sinks even below this.

We learn from this that **water contracts when it loses heat.**

Experiment 75. — Place on the two pans of a pair of scales two beakers of the same weight containing exactly the same volumes of hot and cold water, say half a pint at 80° C. and half a pint at 10° C.

Observe that the cold water is heavier than the same bulk of hot water.

We learn from this that equal volumes of water at different temperatures have not the same weight.

Experiment 76.—Suspend two thermometers, *a*, *b*, from a retort stand, as shown in Fig. 40, and allow them to dip into a large beaker of cold water. Place the beaker in a pail containing a freezing mixture of snow and salt or salt and powdered ice.

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Observe that the temperatures recorded by *a* and *b* fall until they both stand at 4° C. After this point is reached the column of mercury in *b* remains at 4° C., whilst that in *a* sinks lower and lower until it reaches 0° C., at which temperature a film of ice is formed upon the surface of the water.

We learn from this that water at 4° C. is heavier than water at lower temperatures, and that this lighter and colder water floats upon the surface until the freezing temperature is reached, at which point it passes into solid ice.

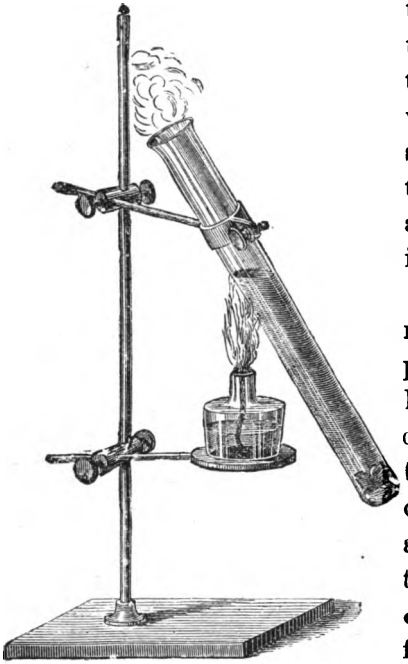

Fig. 41 —Water a bad conductor of heat.

After this last experiment the student may perhaps ask the question, How is it that the heat does not quickly travel through the water from one part of the mass to another? The answer is that **water** is a **bad conductor of heat,** and this fact may be illustrated by the following :—

Experiment 77.—Place a few pieces of ice at the bottom of a long test-tube, Fig. 41, and three parts fill the test-tube with water. Now heat the upper portion by means of a spirit lamp.

Observe that the water at the top grows hot, and may even be boiled, and yet the ice remains unmelted.

We learn from this again that hot water is lighter than cold water, for it floats at the surface ; and that water is a bad conductor of heat, for the ice remains unmelted.

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Experiment 78.—Close a short piece of glass tubing at one end, and draw the other end out to a point. Place the open end of the tube in water and heat the other end: the air will be driven out through the water, and as the tube cools water will rise to partially fill it. Boil the water in the tube and thus expel the air dissolved in it; whilst it is boiling place the open end in cold water, the steam condenses, and more water rises to completely fill the tube. Now seal up the drawn-out end in the blowpipe flame. Place the tube which is now quite full of water in a freezing mixture.

Observe that the water freezes and the tube is broken.

We learn from this that water expands when it passes into ice, and is therefore volume for volume lighter than water.

Formation of Currents in Water by Heating.—We have seen that water is a bad conductor of heat, in experiment 77, and that hot water is lighter than cold; and in our second lesson we saw that in liquids the force of cohesion is so slight that the particles of the body glide freely about each other. For these reasons, when water is heated currents are formed; the existence of these may be well demonstrated by the following experiment.

Experiment 79.—About half fill a large flask with water, place it on a sand-bath and apply heat. Now add a little bran or heavy sawdust.

Observe that currents are established in the water, the movements of which are indicated by the sawdust.

We learn from this that **convection currents** are produced in water by heating.

The water at the bottom of the flask becomes heated, and therefore lighter bulk for bulk than the surrounding mass. It in consequence rises through the colder liquid, and as fast as it ascends cooler, heavier water sinks into its place, thus keeping up a constant circulation in the mass, successive portions of the water coming near to the source of heat, and then flowing away until the boiling point is reached. This

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process of heating by the circulation of the mass is termed **convection.** On the earth's surface the great **convection currents** in water give rise to **ocean currents.**

The Boiling Point of Water.—We have seen that the application of heat to water causes a rise in temperature and a change in volume ; but this rise in temperature does not go on indefinitely, for a point is ultimately reached beyond which the temperature does not rise, no matter how much heat is applied to the water. The water at this stage is said to boil, and the temperature recorded by the thermometer is called the *boiling point.* When cold water is first heated the air dissolved in it expands and escapes ; then portions of the liquid in contact with the heated surface expand and rise, and colder portions flow in to take their place. These convection currents continue, and presently a point is reached at which the heated surface converts the water in contact with it into steam, and this steam is condensed by the colder water around. But after a time the whole mass of the water reaches a temperature of 100° C. ; then bubbles of steam, formed at the bottom of the vessel, pass through the hot water and escape into the air. This effect is only produced when the pressure of the vapour within the bubbles of steam is able to overcome the pressure of the air. The temperature at which water boils will depend therefore upon the pressure to which it is subjected.

Experiment 80.—About half fill a flask, provided with a long neck, with water, and boil the water. When the liquid has been boiling for a short time, quickly remove the flame and well cork the flask. Next invert the flask, and squeeze from out a sponge, or pour over it, some cold water.

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Observe that when the source of heat is removed, the water stops boiling; but on pouring cold water over the inverted flask, the liquid in the flask commence to boil again.

We learn from this that the **boiling point depends upon pressure**; for the steam which escapes during boiling drives out the air, and when the steam is condensed, the pressure within the flask is so reduced that the water boils again at once.

Since the boiling point depends upon atmospheric pressure, we find that in a mine water will boil at a higher temperature than at the sea-level; whilst on a mountain water boils at a lower temperature than at the sea-level.

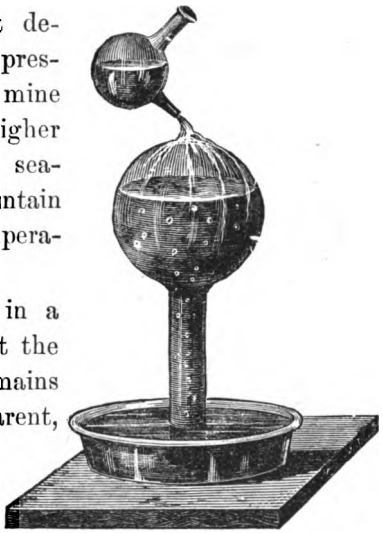

Fig. 42.—Boiling point depends on pressure.

When water is boiled in a glass flask we observe that the flask above the water remains quite clear and transparent, while a dense white cloud floats round the mouth of the vessel. Near the spout of a kettle a white cloud plays when the water is boiling, and between this cloud and the spout no visible gas or vapour exists. In each of these cases the **invisible gas is steam**, and when this gas passes out into the cool air it is condensed into a **cloud of small water particles**, which is sometimes called steam, but is really a cloud of **condensed vapour**, for steam, as we have just seen, is quite invisible.

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Experiment 81.—Place a small kettle of water on a tripod stand, heat it by means of a Bunsen burner, and in the cloud of vapour hold the flame of a spirit lamp.

Observe that after a time the water boils, steam is given off and condenses to form a cloud of vapour just beyond the spout. This vapour passes into an invisible gas on the application of heat.

We learn from this that steam is invisible, and that a cloud of condensed vapour may be made invisible by the application of heat.

Experiment 82.—Half fill a flask with cold water. Provide a sound cork with two holes; through the larger of these pass a thermometer, and through the other an open piece of glass tubing bent at right angles. The thermometer should not touch the liquid, but be so placed that the steam perfectly surrounds it when the water boils.

Fig. 43.—Steam is invisible.

Observe that the thermometer records 100°. when the water boils, and that it remains constant while the water continues to boil.

Repeat this experiment, using alcohol, and then solutions of solids and liquids in water; for example, a mixture of alcohol and water, or a solution of washing soda in water, and observe that the boiling point is different in each case.

We learn from this that the boiling point of water is **a constant** at any given pressure, and that water with solid matter in it has a higher boiling point than pure water. For this reason sea water boils at a higher temperature than fresh water.

All liquids have their own boiling points; thus alcohol

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and ether boil at lower temperatures than water, the former at 78° C., and the latter at 55·5° C. A mixture of alcohol and water therefore boils at a temperature between 78° C. and 100° C. Turpentine and oils have higher boiling points than water; the former boils at about 158° C.

When **water is converted into steam,** or any other liquid is changed into a gas or vapour, it **undergoes an enormous increase of volume.** If we completely boil away one pint of water into steam and then measure the

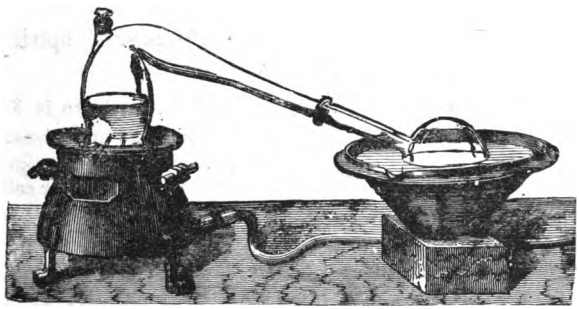

Fig. 44.—Distillation of water.

whole of the vapour produced from it, we should find that one pint of water would yield 1700 pints; so that steam occupies 1700 times the volume of the water from which it is formed.

Distillation.—When water is converted into steam, any solid saline impurities which it may contain are left behind as we saw in experiment 32. The liquid which is driven off as a vapour in the act of **evaporation** may be recovered by cooling it. In this way it is made to undergo **condensation,** passing once again into the liquid form. By such a process fresh water may be obtained from salt water,

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and in nature the operation is carried on upon a gigantic scale in the formation of rain, dew, etc.

Water prepared from steam in this manner is known as **distilled water**, and the apparatus which is employed for the distillation of water or any other liquid is called a **still**. Distilled water is chemically the same as ordinary water, but if the operation has been carefully performed it is freed by this means of all impurities held in solution that are not volatile at temperatures below 100° C.

The process of distillation is employed on a large scale commercially in the preparation of intoxicating spirits and in the manufacture of coal gas.

Experiment 83.—Fit up a retort and flask, as shown in Fig. 44. Place in the retort a solution of sulphate of copper. Now apply heat.

Observe that the solution boils, and that the steam condenses in the flask to a clear, colourless liquid. After a time examine the water collected, and notice that it is without any decided flavour, although somewhat flat to the taste.

We learn from this that it is possible to obtain fresh, clear water from such a solution by the process of distillation, and that this operation involves the evaporation of the liquid through the agency of heat and the condensation of the vapour by the withdrawal of heat.

Experiment 84.—Pour some water prepared in last experiment into a watch glass, and heat it on a sand bath.

Observe that after all the liquid has disappeared the glass is quite clean, no residue remains. Repeat the operation, using, instead of distilled water, tap water, and observe that in this case a most decided residue is left in the glass.

We learn from these experiments that **distilled water contains no solid matter** in solution, but that ordinary drinking water is charged with dissolved salts which are left behind when the water is evaporated.

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In Fig. 45 we have a **worm condenser** and **still,** by means of which large quantities of impure water or other liquids may be purified by distillation.

We have seen that the body produced from water by boiling is truly **gaseous water,** for it may be at once converted into the liquid by cooling it; and it is identical in

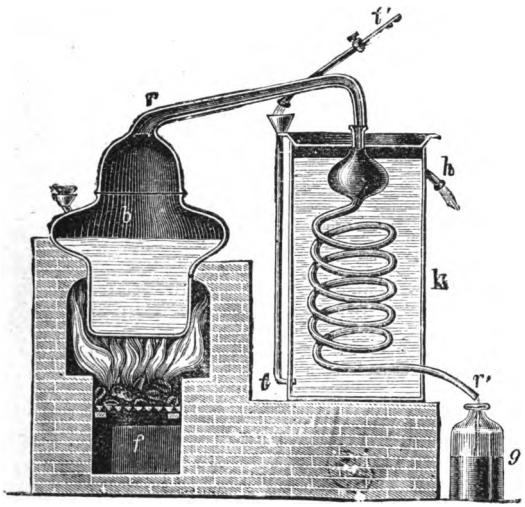

Fig. 45.—The worm condenser.

all respects with that water which is obtained by melting pure ice. Chemically, therefore, these bodies—ice, water, steam—are the same. Now if the steam obtained by boiling water is passed over red-hot iron, it undergoes a most marked change, for the resulting body is not steam, nor is it in any sense like it; for it cannot be condensed into water, and on the application of heat it burns with an almost colourless flame.

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Experiment 85.—Fit up a piece of iron gas tubing or glass combustion tubing, with a delivery tube at one end opening beneath the surface of water in a pneumatic trough, and at the other ending in a small flask partially filled with water. Place in the tube some coarse iron filings or small nails, and then fix it in a furnace or place beneath it two large fish-tail Bunsen burners. When the tube is quite hot, boil the water, the steam will then pass through the tube over the red-hot nails.

Observe that an invisible gas collects in the gas jar, which will burn, and that the combustion tube at the close of the experiment no longer contains bright iron filings, but only iron rust.

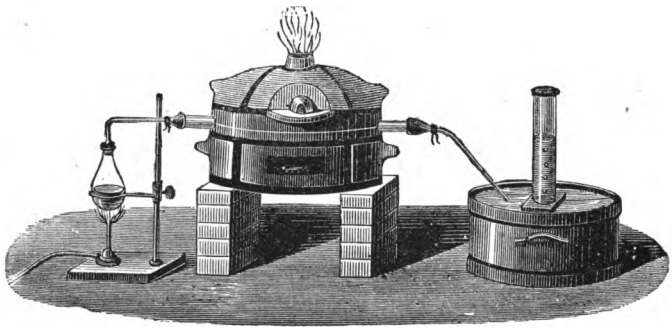

Fig. 46.—Decomposition of water by red-hot iron.

We learn from this that **red-hot iron will decompose gaseous water** with the liberation of an invisible combustible gas ; and that the steam also yields a body which combines with the heated iron and causes it to become rusty. The combustible gas obtained by this means from water is known as **hydrogen,** and that which causes the rusting of the iron is termed **oxygen.** Thus, then, we see that **water is a chemical compound,** for it can be split up into two very different substances. By means of an electric current these two bodies may both be isolated from water as gases.

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SUMMARY TO CHAPTER VI.

WATER

- **Occurrence in Nature.** — In the free state it covers a very large portion of the earth's surface; it enters into the composition of many minerals, and is an essential constituent of all plants and animals. It is never met with pure in nature, but always holds gases and solids in solution.
- **Physical Properties.**
  - Exists in three states—ice, water, steam.
  - Expands when heated above 4° C.
  - Expands when cooled below 4° C.
  - Maximum density reached at 4° C.
  - Increases in volume when it passes into ice.
  - Bad conductor of heat. Heated by convection.
  - At the sea-level it boils at 100° C. Steam invisible.
  - Freezes at 0° C. Ice floats on water, because it is bulk for bulk lighter than water.
  - Boiling point depends upon the pressure to which the liquid is subjected and the purity of the water.
  - All liquids have special boiling points.
- **Purification of Water.** — Distillation. Conversion of a liquid into a vapour and condensation of same. Thus solids in solution are separated from their solvents.
- **Chemical Composition.** — Water is a chemical compound of oxygen and hydrogen.

QUESTIONS ON CHAPTER VI.

1. What are the physical conditions in which water is known? Give examples.

2. What do you understand by the maximum density of water?

3. How would you prove that heat has an effect upon the bulk of water?

4. Describe how you would prove that water expands when it freezes.

5. Why does ice form on the surface of water and not at the bottom?

6. How would you demonstrate that water is a bad conductor of heat?

7. What are convection currents? How would you demonstrate that they exist in water which is being heated?

8. Show how you would prove that steam is invisible.

9. How is the boiling point of a liquid determined, and how would you show that the boiling point depends upon pressure?

10. How is it that boiling water is not very hot on the top of a high mountain?

11. Show fully how you would obtain fresh water from salt water.

12. How may it be proved that water is a chemical compound?

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Preparation for Seventh Lesson.

APPARATUS AND REAGENTS REQUIRED.

Combustion tubing, ordinary glass tubing, iron turnings, corks, cork borers, balance, flask, water, granulated zinc, Woulffe's bottle, pneumatic trough, gas jars and plates, dilute sulphuric acid, balloon, copper oxide, apparatus as in Figs. 50 and 51, eudiometer, mercury, voltameter, electric battery.

EXPERIMENTS TO BE PERFORMED.

"To show the presence of hydrogen in water, pass steam through a red-hot iron tube filled with coarse iron turnings or nails. Note the change in the iron both in appearance and weight. This increase of weight and the weight of gas which comes off equals weight of steam which has been decomposed. Hence two substances in water: one the combustible gas that comes through the tube, the other the body which remains with the iron. Collect the hydrogen over cold water in proof that it is not steam; also show that it burns.

Explode a mixture of two volumes of hydrogen and one volume of oxygen.

Plunge a burning taper into a jar of hydrogen held mouth downwards, to show burning of the gas and extinction of the taper.

Show by a balloon or soap bubbles, or inverted beaker glass suspended from a balance, that hydrogen is lighter than air.

Condense the water formed by the burning of a jet of hydrogen."—*Science Department's Syllabus.*

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CHAPTER VII.

CHEMICAL PROPERTIES OF WATER.

Hydrogen and its Properties—The Burning of Hydrogen in Air, and the weight of the Product (Water) compared with the weight of the Hydrogen—The Difference due to the Oxygen of the Air with which Hydrogen has combined—Hence Oxygen and Hydrogen are the Constituents of Water — Combination of Oxygen and Hydrogen with Explosion to form Water—If by measure there be twice as much Hydrogen as Oxygen, or by weight eight times as much Oxygen as Hydrogen, then no Gas remains: all becomes Water—All Water composed of these two Bodies in this Proportion—These two Bodies can then be separated from Water and can be made to unite and form Water—In all cases of Chemical combination Bodies are united in constant Proportion.

WATER, we learnt in last lesson, from experiment 85, **is a chemical compound** which may be decomposed by the action of red-hot metallic iron. By this means a gas was obtained from it, in all respects quite unlike gaseous water or steam, with the exception that it is invisible.

Experiment 86.—Place some iron turnings in a short piece of combustion tubing, heat the tube in order that the contents may be made thoroughly dry, then provide it with a sound cork at each end. Carefully note the weight of the tube with its contents. Fit into the corks delivery tubes, one of which should dip into water, and the other pass into a small flask in which steam may be generated. When the apparatus is fitted, raise the combustion tubing and its contents to a red-hot state, and then, maintaining the heated condition of the iron, boil the water in the flask, and thus drive steam over red-hot iron filings. Collect the gas which comes over. After the action has gone on for some time allow the apparatus to cool, and weigh the combustion tube and its contents again.

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Observe the change in the appearance of the iron and the increase in weight of the tube. Examine the collected gas, and notice that it will not support combustion, but will burn.

We learn from this that two substances may be obtained from water — one, the combustible gas that is collected in the jar cannot be steam, for if it were it would be condensed by the cold water through which it is made to pass ; this body is **hydrogen.** The increase of weight of the iron is due to the body which has combined with it ; this is known as **oxygen.**

If all the gas which comes off be collected and weighed, and the weight added to the increase in weight of the tube containing the rusted iron, the total weight will exactly equal the weight of the steam which has been decomposed.

The method of preparation and properties of oxygen we studied in Chapter V., after we discovered that it was the active constituent of air. Since the combustible gas hydrogen is a new body, the student should prepare a larger quantity than can be easily obtained by the above plan, in order that the properties of the substance may be more fully studied.

HYDROGEN, as prepared in experiment 86, is a colourless, inodorous, tasteless gas. It occurs free in nature in small quantities mixed with the gases emitted by volcanoes. Mainly, it is met with chemically united with oxygen in water, of which it forms nearly one-ninth by weight. Combined with carbon, nitrogen, and oxygen, it occurs in living things. It is a constituent of coal, vinegar, alcohol, turpentine, fat, starch, flour, soap, sugar, and gum. Hydrogen is the lightest known substance, being about fourteen

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and a half times lighter than air. On account of this extreme lightness it is sometimes employed for filling balloons. It is inflammable, and burns with an almost colourless flame, but a lighted taper is extinguished by it. Being slightly soluble in water, and much lighter than air, it may be collected in jars by upward displacement. Since hydrogen mixed with air or oxygen yields a body which explodes at once when a flame is brought near, it is not advisable for the young student to collect it by displacement, and in all cases care should be taken to ascertain that the air has been driven out of the apparatus used for its preparation.

When hydrogen burns in air or oxygen, water is produced; and if all the water yielded by burning a certain weight of hydrogen be carefully collected we shall obtain nine times as much water by weight as hydrogen is employed. Eight-ninths of the resulting compound, water, consist of oxygen.

Preparation.—This gas may be prepared by the action of many metals upon water. In any case the **metal fixes the oxygen** of the water by chemically uniting with it, and **liberates**, or sets free, **the hydrogen.**

Experiment 87.—Fit up a 4-oz. flask or *Woulffe's bottle* with *corks, thistle tube,* and *delivery tube,* as shown in Fig. 47. See that the thistle tube reaches nearly to the bottom of the bottle, and that the delivery tube passes just through the cork. Remove one of the corks and place some pieces of *granulated zinc* in the bottle, adjust the cork, attach a piece of indiarubber tubing to the end of the delivery tube, and allow the same to dip into a trough of water, as shown in Fig. 47. Now pour down the thistle tube enough water to cover the zinc, and then add dilute *sulphuric acid.*

Observe that effervescence occurs, and that bubbles of a colourless gas escape through the water. A little of the gas should be allowed to escape for a few minutes in order that all the air may be driven out, otherwise an

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explosive mixture will be produced. A small quantity of the gas collected in a test-tube should burn at the mouth of the tube with a quiet flame. It is then safe to collect the gas. Collect two or three jars of this gas.

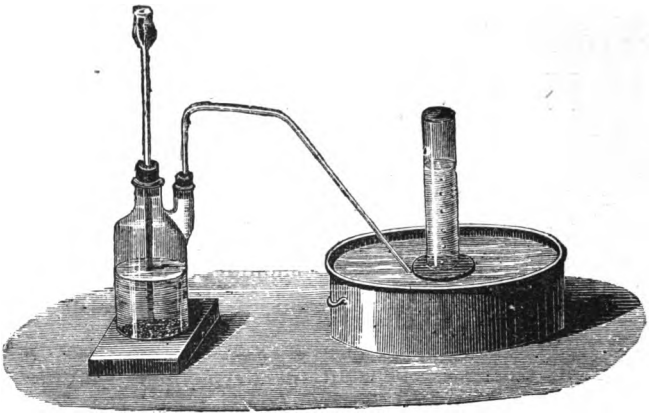

Fig. 47.—Preparation of hydrogen.

We learn from this that hydrogen may be prepared by the action of dilute sulphuric acid upon the metal zinc.

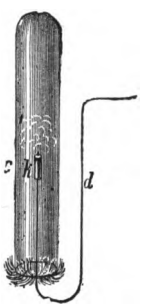

Fig. 48.—Combustion of hydrogen in contact with air.

Experiment 88.—Hold a jar of the gas *mouth downwards*, and apply a lighted taper; then plunge the burning body into the jar.

Observe that the gas burns at the mouth of the jar, where it is in contact with the air, with a quiet flame, and that the flame is extinguished at once when the burning body is immersed in the gas; but that it is rekindled by the burning gas when it is withdrawn.

We learn from this that hydrogen will burn in air, but that it will not support combustion.

Experiment 89.—To one arm of a balance, as shown in Fig. 49, attach a gas jar or light beaker containing air, with its *mouth downwards*, and carefully counterpoise it with shot. Then introduce into the jar the delivery tube from which hydrogen is escaping; the light gas will rise in the jar and drive out the air.

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Observe that as the air is displaced by the hydrogen the jar rises, and that as the air passes back again the vessel increases in weight.

We learn from this that hydrogen is lighter than air. This fact may be demonstrated also by filling a balloon, or by blowing soap bubbles with hydrogen.

Another convenient method of proving that hydrogen is lighter than air, in the absence of a balance, is to place two jars mouth to mouth, as shown in Fig. 12. The lower

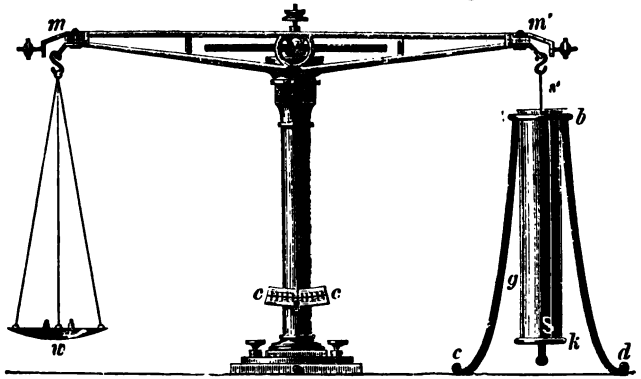

Fig. 49.—Weighed jar of air decreasing in weight with admission of hydrogen.

one should be filled with hydrogen, and capped with a well-greased ground glass plate, and the upper one should contain air; when the jars are placed well in position one over the other, the plate should be removed. On examination it will be found that the upper vessel now contains hydrogen and the lower one air. When the jars are kept long in position with open mouths the gases freely diffuse, and on examination both will be found to contain explosive mixtures of air and hydrogen.

Experiment 90.—Fit to the delivery tube of the hydrogen apparatus a drying tube containing some pieces of *dry calcium chloride*, and attach to the end a piece of glass tubing drawn out to a jet; in the open

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end of the tube a small piece of platinum foil in roll form should be fused to protect the glass. The hydrogen in passing through the tube is thoroughly dried, and when a flame is brought to the jet the gas burns with a quiet, colourless flame. Hold over the burning gas a dry, clean, cold bell jar.

Observe that drops of water collect and run down the glass.

We learn from this that dry hydrogen burnt in air produces water. If the flame decreases, a little more acid should be added to quicken the formation of the gas.

The word hydrogen is derived from *hudor*, water, and *genesis*, a generator.

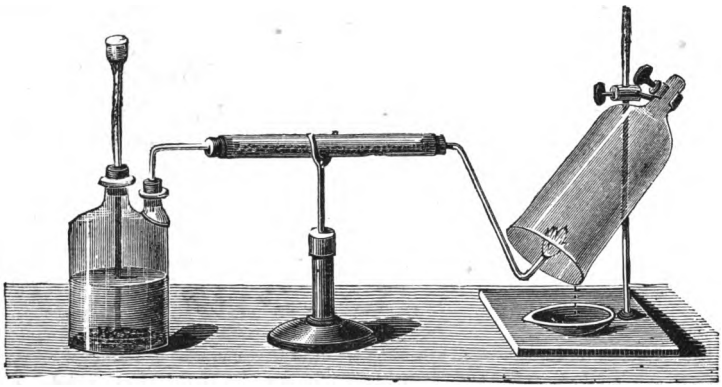

Fig. 50.—Dry hydrogen burnt in air produces water.

Experiment 91.—Put out the flame produced in experiment 90, and attach to the end of the delivery tube, from which the dry hydrogen is escaping, a light balloon, and allow the gas to fill it; when full liberate it.

Observe that the balloon of dry hydrogen rises quickly through the air. If the balloon contains some air or oxygen before the hydrogen is passed into it, on the application of a flame it will explode violently.

We learn from this again that hydrogen is lighter than air, and that with this body or oxygen it forms an explosive mixture. In all experiments, therefore, in which a flame has to be brought near to apparatus containing hydrogen, care should be taken to ascertain that the whole of the air has been expelled, otherwise dangerous explosions may result.

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Hydrogen is a Reducing Agent.—Not only will hydrogen, when heated with the free oxygen of the air, readily form water, but it possesses such a strong chemical affinity for oxygen that it will also tear it away from many compounds in which it is united with metals. When hydrogen so acts upon metallic oxides it **reduces** them to the metallic state. In manufacturing and analytical operations hydrogen and bodies rich in hydrogen are therefore frequently employed in the chemical process of **reduction.**

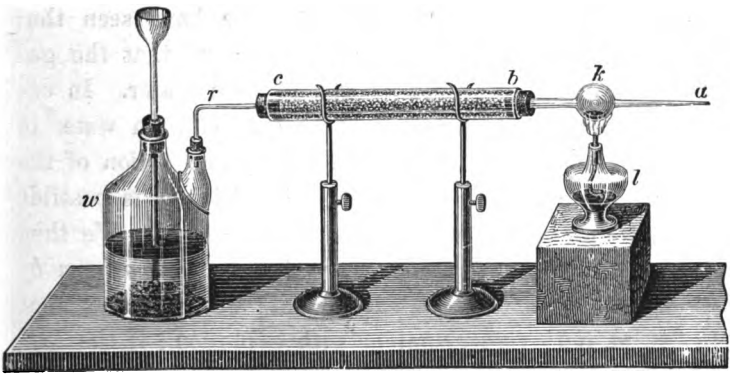

Fig. 51.—Reduction of copper oxide by hydrogen gas.

Experiment 92.—In Fig. 51 is shown a piece of apparatus fitted up to demonstrate this property of hydrogen. The gas is generated in the bottle *w*, and dried by the action of calcium chloride in the tube *c b*, the bulb *k* is charged with *copper oxide*, which is heated by means of a Bunsen burner or spirit lamp at *l*.

Observe that after the gas has escaped for a few seconds, and when the copper oxide is red hot, steam escapes from the end of the tube *a*, and the oxide of copper is reduced to the bright-red metallic condition.

We learn from this that hydrogen acts as a reducing agent, for it will decompose copper oxide with liberation of metallic copper and formation of steam.

The flame of burning hydrogen is only very slightly

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luminous as we have seen, but it is extremely hot, and, if oxygen gas is driven into it, the heat becomes so intense that few substances can resist its enormous melting power. Silver, gold, and platinum melt in the flame thus produced. Lime, subjected to the heat of the **oxy-hydrogen** flame, becomes white hot, and gives out the intense brilliant white light known as the **lime-light.** The lime does not melt, but is raised to the glowing condition under the influence of the great heat.

COMPOSITION OF WATER.—We have seen that hydrogen may be obtained from water, and that the gas combines chemically with oxygen to form water. In experiment 86, we did not obtain the oxygen from water in the form of a gas, for it was fixed by the action of the red-hot iron; but by employing electricity it is possible to break up water into its two component gases. We then find that there is *twice as much hydrogen as oxygen by volume in water.* In our experiments we have seen that hydrogen and oxygen combine with explosion; if the *bulk* of the hydrogen is twice that of the oxygen, or by *weight* eight times as much oxygen as hydrogen, no gas remains—all becomes water. And water, under all circumstances, is composed of these two bodies in this proportion.

Chemists are able to demonstrate the exact composition of water by various methods, all of which are based upon the facts of our past experiments. The methods they employ are **analytical** and **synthetical,** and they may be **volumetric** or **gravimetric.** In the volumetric analysis of a body, we employ methods by which we can determine *the proportion by bulk,* in which gases combine to form

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a given compound. By the gravimetric method, we are able to discover *the proportion by weight* in which elements combine to form any substance.

Volumetric Analysis.—The instrument made use of is known as a **voltameter,** a simple form of which may be constructed in the following manner: Cut about two inches off the stem of a large glass funnel; into the remaining short piece of the stem fix an india-rubber stopper, through which two platinum wires should be passed, to which small strips of platinum foil must be attached. Fill the glass vessel thus formed with water, slightly vcidulated with sulphuric acid to make it a better conductor of electricity; two large test-tubes should now be filled with water, and carefully inverted, one over each of the platinum plates. If the two wires which pass outward through the indiarubber stopper be connected with the terminals of an electric battery, a brisk evolution of gas takes place. The bubbles ascend from the platinum plates through the water, and displace the liquid in the tubes. After the decomposition has gone on for some time, one tube will be found to contain twice as much gas as the other; if the tubes be removed and the gases tested, the contents of the one containing two volumes will be found to possess the

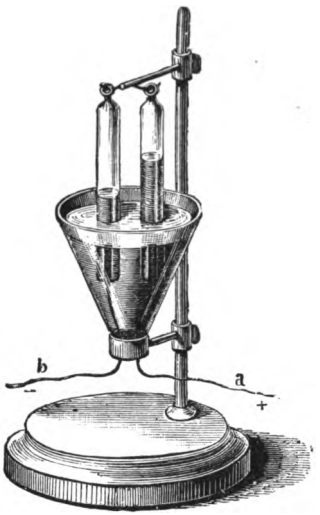

Fig. 52.—Analysis of water by electricity.

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properties of hydrogen, whilst the one volume of gas in the other tube will be found on examination to exhibit the characteristics of oxygen.

A more convenient form of apparatus devised by Dr. Hofmann is shown in Fig. 53. It consists of a U tube connected in its bend with a third tube on which a bulb is blown at *A*. The two limbs of the U tube end in jets provided with stopcocks *B*, *C*. These should be filled with water acidulated with sulphuric acid. At the lower portion of the U tube, and near the bend of the tube, two platinum wires pass through the glass and end on the inside in plates of the same metal. The main tube *A* is capacious enough to receive the liquid contained in both limbs of the U tube, and it acts as a reservoir. When gases are liberated in *B*, *C*, the liquid is forced back into *A*, and when the stopcocks are turned and the jets opened, the pressure of the liquid is sufficient to drive out any gas which has been collected in *B*, *C*. As soon as the platinum wires are connected with a battery of four or five Bunsen cells, bubbles of gases are seen to form on the platinum plates, from which they are liberated, and rising through the water are collected in *B* and *C*.

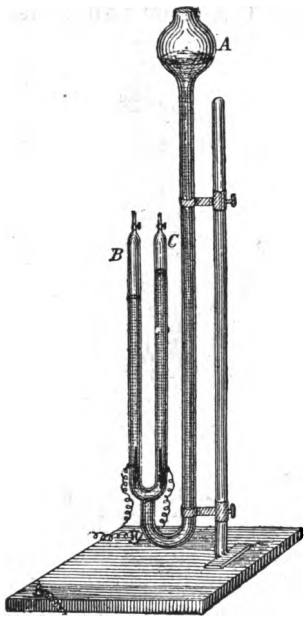

Fig. 53.—Hofmann's voltameter.

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On examination they prove to be colourless, and the quantity collected in one tube is about twice that collected in the other. The gases may be tested by opening the stopcocks and applying successively a glowing cedar splint and a lighted taper. If the cedar splint bursts into flame, the gas is oxygen ; if the lighted taper ignites the gas, it is hydrogen. Under all conditions when this experiment is repeated, we observe invariably the same **constant volume proportions,** and we therefore infer that the two gases are combined in water in these proportions, thus proving that water yields hydrogen and oxygen gases, in the proportion of two volumes of the former to one volume of the latter.

Fig. 54.—Charging a eudiometer tube.

Synthesis.—Hydrogen and oxygen if mixed will combine together with great force when heat is applied, often giving rise to a dangerous explosion. If a strong lemonade or soda-water bottle be filled with a mixture of hydrogen and oxygen in the proportion of two-thirds of the former to one-third of the latter, and a light applied at the neck of the bottle, a violent explosion follows, and a thin film of vapour is deposited on the sides of the vessel. It is advisable to adopt the precaution of wrapping the bottle in a cloth before the gases are exploded.

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For accurate synthesis of water a long, graduated, strong glass tube closed at one end, and known as a **eudiometer,** is taken. Two platinum wires *a*, *b*, are melted into the upper and closed part of the tube. The tube is filled with mercury, and then inverted in a vessel, *g*, containing the same liquid metal. Hydrogen is now introduced and allowed to displace two volumes of mercury, after which one volume more is displaced by oxygen. We have a tube containing two volumes of hydrogen and one volume of oxygen. This is next fixed so that the lower or open end is tightly pressed against a pad of india-rubber at the bottom of the trough containing the mercury. An electric spark is next made to pass through the mixed gases by means of the platinum wires in the upper part of the tube; an explosion follows, and the water thus produced is deposited as a thin film upon the inner surface of the tube. On raising the tube from the india-rubber pad, the mercury rushes up to take the place formerly occupied by the mixed gases; for the water produced by the combination of the gases only occupies about $\frac{1}{2000}$ of the bulk.

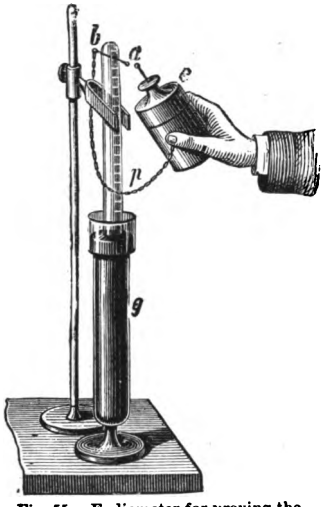

Fig. 55.—Eudiometer for proving the composition of water synthetically.

These and many other experiments have taught chemists that in all cases of chemical combination bodies are united in constant proportion.

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SUMMARY TO CHAPTER VII.

WATER.

- **Composition** — Water consists of hydrogen and oxygen united in the proportion of two parts to one part by *volume* and one part to eight parts by *weight*. Proved by *analysis* and *synthesis*, volumetrically and gravimetrically.

HYDROGEN

- **Occurrence in Nature** — Chiefly in water, and combined with other elements in organic substances.
- **Properties** — Lightest known body, invisible, inodorous, tasteless, will not support combustion, but will burn in air or oxygen to form water. Behaves as a *reducing* agent.
- **Preparation** — May be prepared by the action of water on metals, which possess the power of fixing the oxygen and liberating the hydrogen. Some metals fix oxygen of water when cold, like sodium and potassium, others must be heated. Steam decomposed by red-hot iron, water decomposed by zinc in the presence of an acid.

QUESTIONS ON CHAPTER VII.

1. Sketch the apparatus you would require for the preparation of hydrogen, and name all the parts.

2. What are the more important properties of hydrogen, and how would you demonstrate the same?

3. How does hydrogen usually occur in nature? Give examples to illustrate your answer.

4. I burn a jet of hydrogen in air; what chemical effect is produced? I plunge the burning jet into oxygen; what chemical change takes place?

5. I have two cylindrical jars of hydrogen; I hold one of them mouth upwards and the other mouth downwards. At the expiration of half a minute I plunge a lighted taper into each jar; describe exactly what you would expect to take place in each case.

6. What is the origin of the word hydrogen? Justify the use of the name.

7. Explain fully how you would prepare hydrogen from steam by the action of red-hot iron.

8. How is the lime-light produced?

9. Explain the following terms—*reducing agent*, *reduction*, *reduced*, as employed by chemists.

10. Describe the apparatus (exclusive of the battery) required to electrolyse water acidulated with sulphuric acid, and make a sketch of it. What are the names and relative volumes of the gases given off, and how would you demonstrate their characteristic properties?

11. Describe an experiment for proving that water is a compound body, and state precisely the conclusions to which the experiment leads you.

12. Describe shortly the experiments to illustrate the analysis and synthesis of water. Sketch the apparatus employed.

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Preparation for Eighth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of the various forms of charcoal and fat, oil, soap, starch, sugar, flour, bread, meat, amber, and jet, a dozen hard glass tubes, strong sulphuric acid, hydrochloric acid, marble, wood, hard glass test-tube fitted with cork and delivery tube, acetic acid, sodic hydrate, crucible and tripod stand, apparatus as fitted in Fig. 61, tapers and lime-water, apparatus as in Fig. 63, litmus papers, apparatus as shown in Figs. 65 and 66, and described fully in pages 127, 128.

EXPERIMENTS TO BE PERFORMED.

"Make charcoal by heating wood in a test-tube provided with a cork and delivery tube.

The blacklead of a pencil as a specimen of graphite.

Show that acids and alkalies do not change charcoal, but that when heated it soon burns away, and only ash is left.

Take a small piece of charcoal in a glass tube, pass air over it into lime-water, and show that no change takes place until the charcoal is made red hot ; as the charcoal disappears the lime-water becomes milky.

Prepare carbonic acid gas by acting on marble with dilute acid.

Show by means of the balance, or by soap bubbles, or by passing it from one vessel to another, that carbon dioxide is heavier than air, that it acts on lime-water, that a burning candle is extinguished in it.

Its solubility in water shown by agitating a tube of the gas over water.

Collect all the gas given off from a small piece of marble weighing 1 or 2 grammes.

Show by collecting in inverted beaker the products of combustion of a candle, of a lamp, and of a gas flame, and adding lime-water, that carbon dioxide is given off.

Show also by means of lime-water that respired air contains this gas."—*Science Department's Syllabus.*

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CHAPTER VIII.

CARBON AND ITS OXIDES.

*Charcoal, Plumbago, Graphite, or Black Lead and Diamond—When Wood, Sugar, Meat, or Bread are heated Carbon remains—Charcoal not changed in the Air at ordinary Temperatures—Combination of Carbon with the Oxygen of the Air at a red heat—Carbon Dioxide, a compound of Carbon and Oxygen—Chemical combination of Carbon and Oxygen is attended by the evolution of a definite amount of Heat expressed by amount of Water it will heat—Combustion—The properties of Carbon Dioxide—Water dissolves Carbon Dioxide at ordinary Temperatures—Action of Carbon Dioxide on Lime-water; no Animal can live in this Gas—*100 *parts of Carbon Dioxide are composed of* 27·27 *parts of Carbon and* 72·73 *parts of Oxygen—Carbon Dioxide obtained from Marble, Lime-stone, Oyster Shells, Chalk, etc.—Charcoal Fire—Coal composed of Carbon, Hydrogen, and a little Oxygen, etc.; its burning is the result of the Carbon and the Hydrogen combining with the Oxygen—Whenever Oil, Tallow, or Coal Gas is burnt Carbon Dioxide and Oxide of Hydrogen (Water) are formed—Respiration produces similar changes—In expired Air the same Products arise as from the burning of the Food, and there is the same evolution of Heat—Carbon a Constituent of all Animal and Vegetable Bodies.*

Carbon is a chemical element of very great importance and interest, for it forms by far the larger proportion by weight of all dried animal and vegetable substances. It is a solid body which is **known in three forms,** which differ so much from each other that unless we had very distinct proof that they are chemically the same we should regard

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them as different bodies, and not three forms of one and the same substance. The three forms are :—

1. Graphite, plumbago, or **blacklead,** a soft, black, somewhat metallic-looking substance which occurs native in the earth, either as crystalline plates or in fine grained masses. This form of carbon is employed in the manufacture of the best blacklead pencils, for lubricating machinery, for crucibles, and for coating metallic surfaces to protect them from the action of the air.

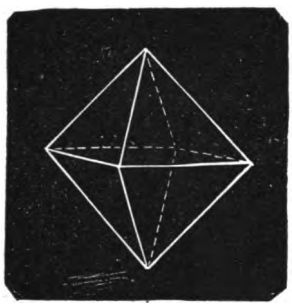

Fig. 56.—Diamond crystal.

2. Diamond, an example of crystalline carbon, is the hardest gem known. This form of carbon is chiefly employed for jewellery, and, from its extreme hardness, it is used for cutting and polishing purposes.

3. Amorphous carbon, or charcoal, occurs nearly pure in lampblack, and in a less pure form in coke, wood charcoal, and animal charcoal. Wood charcoal is very porous, and floats on water in consequence of the air contained within its pores. Amorphous carbon is employed in making printers' ink, for filtering and decolorising purposes.

The different forms of carbon vary very much in physical appearance, colour, hardness and density, but they all *yield equal quantities of carbonic acid gas when the same weights are burnt in oxygen.*

In nature carbon exists in its purest state in diamonds, which are found in limited numbers in rocks in certain

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districts ; in a less pure state it occurs combined with other elements as coal. In the atmosphere, in small quantities, carbon is found combined with oxygen as carbonic acid gas ; combined with lime, and magnesia and iron, it is very abundant in rocks. Combined with hydrogen it forms turpentine, coal gas, paraffin, and benzoline ; and chemically united with hydrogen and oxygen it is present in sugar, gum, oil, wood, cotton, starch, fat, meat, and bone.

Experiment 93.—Place in four hard glass tubes separate pieces of *starch*, *sugar*, *bread*, and *meat*, and strongly heat each.

Observe that gases and water are evolved, and that a residue of charcoal remains in each tube.

We learn from this that when starch, sugar, meat, and bread are strongly heated carbon remains, and water and gases are given off. These organic substances must therefore be rich in carbon, and this fact may be further demonstrated by the following experiment.

Fig. 57.—Separation of carbon from sugar.

Experiment 94.—Grind a small piece of loaf sugar to a fine powder, and then warm it in a test-tube with a small quantity of water, and thus prepare a syrup. Place the tube in a stand, and pour a few drops of strong sulphuric acid into the liquid.

Observe that steam is evolved and that the whole mass becomes black and froths up so as to run over the sides of the test-tube.

We learn from this that sugar contains carbon.

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Experiment 95.—Place a few *pieces of wood* in a hard glass tube provided with a delivery tube, and strongly heat the tube.

Observe that a gas is given off which will burn with a luminous flame, that a dark coloured tarry liquid distils over, and at the end of the experiment a residue of **wood charcoal** remains in the tube.

We learn from this that when wood is heated a combustible gas is evolved and charcoal remains.

From experiments 93, 94, and 95 we learn that the **organic bodies,** starch, sugar, bread, meat, and wood, **are all rich in carbon.**

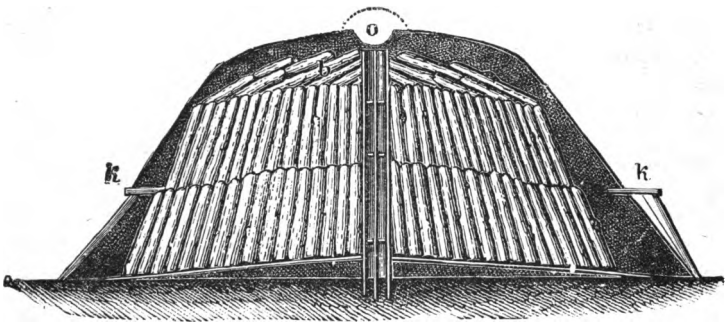

Fig. 58.—Preparation of wood charcoal.

PREPARATION OF CARBON.—This body is prepared in an impure form during the **destructive distillation** of coal in the manufacture of coal gas as **coke.** Other varieties of charcoal are prepared by **roasting** wood or bone respectively.

Vegetable Charcoal is largely employed on the continent as fuel; it is prepared by heaping together faggots of wood, leaving a shaft in the middle, as shown in Fig. 58. The stack is covered with sods of earth, and small inlets are left for air round the bottom of the stack. Burning

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faggots are introduced into the central shaft, and by this means slow combustion of a portion of the wood takes place, whilst the rest of the stack is **carbonized.** **Lampblack** is produced by burning bodies rich in carbon in a limited supply of air. Thus, when oils and resins are burnt, dense clouds of smoke are produced, which may be collected as lampblack.

Fig. 59.—Combustion of carbon on heating.

Charcoal is not changed in the air at ordinary temperatures, and it is an inert substance which is not acted upon by water, ordinary acids, or alkalies. For this reason, woodposts are charred before they are put into the ground in order to prevent their decay ; and blacklead is employed as a coating to prevent the rusting of iron.

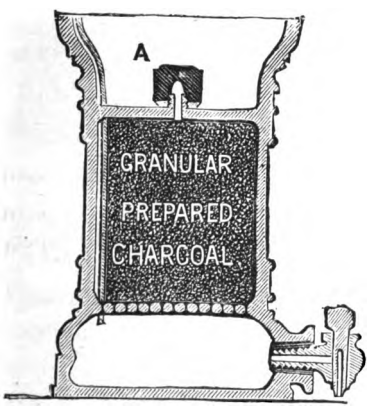

Fig. 60.—Charcoal Filter.

Experiment 96. — Place pieces of charcoal in each of three test-tubes, and to the first add *water*, to the second some *acetic acid*, and to the third a little *sodic hydrate.*

Observe that the charcoal is not acted upon by the water, the acid, or the alkali.

We learn from this that charcoal is insoluble and is a very inert body, exhibiting little chemical affinity for other substances.

Experiment 97.—Place a small quantity of powdered charcoal in a crucible, and heat it in the Bunsen flame, Fig. 59.

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Observe that the carbon slowly burns away, and ultimately leaves only a fine, white ash.

We learn from this that carbon when raised to a red heat burns in air, and that **charcoal is not pure carbon,** for after combustion a white ash remains.

Experiment 98.—Pass a current of oxygen or air from a gasholder through a hard glass tube containing some pieces of carbon, conduct the air into lime-water. Now heat the charcoal until it is red hot.

Observe that the air which passes over the *cold carbon* has no action on lime-water, but that when the charcoal is made red hot it burns, and as

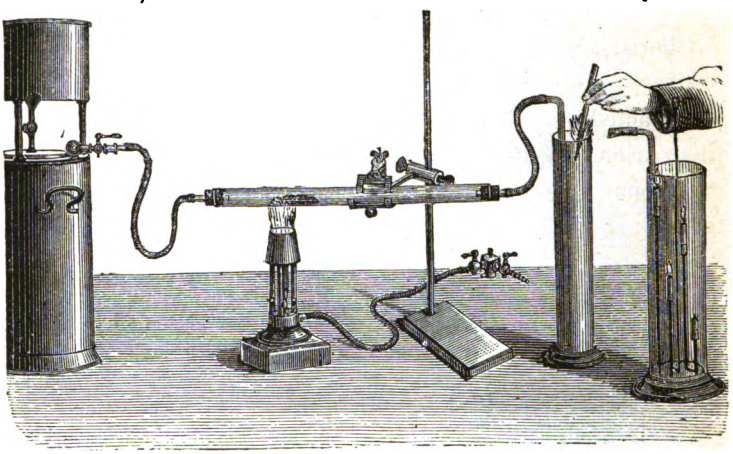

Fig. 61.—Combustion of carbon in oxygen or air.

the carbon disappears the lime-water becomes milky, and that the escaping gas extinguishes a burning taper.

We learn from this that carbon dioxide is a compound of carbon and oxygen, and that it is formed during the combustion of carbon.

The chemical combination of carbon and oxygen is attended by the evolution of a definite amount of heat which is expressed by the amount of water which it will raise through a certain range of temperature.

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COMBUSTION.—The substances which are commonly employed for heating and lighting purposes consist principally of carbon and hydrogen, and it is the chemical combination of these elements with oxygen which leads to the evolution of heat and light. The **products of the complete combustion** of bodies containing hydrogen and carbon, as coal, tallow, paraffin, coal gas, etc., are **carbonic acid gas and water.** Since carbonic acid gas is one of the substances resulting from all ordinary combustions, we may, with advantage, now prove that it is so produced and then study its properties in detail.

Experiment 99.—Place a small quantity of lime-water in three gas jars. Introduce into one a taper attached to a piece of bent wire, in the second place a piece of sponge which has been soaked in alcohol. Fit up a piece of glass tubing, which has been bent so that it may be thrust into the jar, by means of indiarubber tubing with a jet from which coal gas escapes. Ignite the *candle, alcohol,* and *coal gas.*

Observe that in each case the flame soon goes out, and that the lime-water is turned milky.

We learn from this that *tallow, alcohol,* and *coal gas all produce carbonic acid gas* on burning, and that they will not continue to burn in a limited supply of air.

Respiration also leads to the formation of carbonic acid gas and liberation of water, as we have seen in experiment 54. In expired air these oxides of carbon and hydrogen arise from the slow combustion which takes place in all parts of the body, and this oxidation or combustion gives rise to the same evolution of heat as the burning of food materials outside the body. It is the presence of this expired carbonic acid gas, associated with organic matters, which makes the air of crowded places so harmful. By

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ventilation we aim at promoting the rapid removal of the impurities which act so perniciously on health.

Carbonic acid gas is a colourless, inodorous gas, with a slightly acid taste. We have seen that it is produced by the combustion of carbon, and that it exists in air in very small quantities. It is the **choke-damp** of the miner, and it has been detected in small quantities in volcanic gases. During fermentation, and from decaying organic matters, it is given off. Carbonic acid gas is present in expired air, and it escapes in large quantities from lime-kilns during the roasting of chalk or limestone. Many spring waters contain considerable quantities of this gas in solution.

Carbonic acid gas will neither burn nor support combustion; the extinction of a taper is the most frequent means employed in testing for its presence, although many other gases possess this property in common with carbonic acid gas. When ·3 or ·4 parts only are present in 100 parts it renders air irrespirable. In this dilute condition it behaves as a **narcotic poison,** acting upon the nervous system and checking the action of the heart. The ill-effects so often produced by breathing the air in a crowded or badly-ventilated room are due to the presence of this gas, associated with organic expired matters.

Carbonic acid gas has been liquefied and solidified by cold and pressure. The gas is somewhat soluble in water, for a given bulk of water at 15° C. will absorb its own volume of the body.

The gas may be prepared in sufficiently large quantities to demonstrate its more important properties, by the action of acids upon marble, limestone, chalk, oyster-shells, etc.

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Experiment 100.—Place some fragments of oyster-shells or broken marble in a Woulffe's bottle, as shown in Fig. 62. Introduce into the bottle sufficient water to cover the marble, then pour down the thistle funnel *b*, some dilute hydrochloric acid, and collect two jars and a flask of the liberated gas by *downward displacement* or over *warm water*.

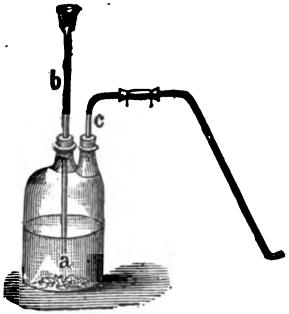

Fig. 62.—Woulffe's bottle for preparation of carbonic acid gas.

Observe that effervescence occurs on the addition of the acid, and that a colourless, inodorous gas escapes through *c*

We learn from this that carbonic acid gas may be prepared by the action of hydrochloric acid upon marble.

Experiment 101.—Bring to the mouth of a jar of carbonic acid gas a burning taper, and then plunge the taper into the jar.

Observe that the gas will not burn, and that the flame is extinguished.

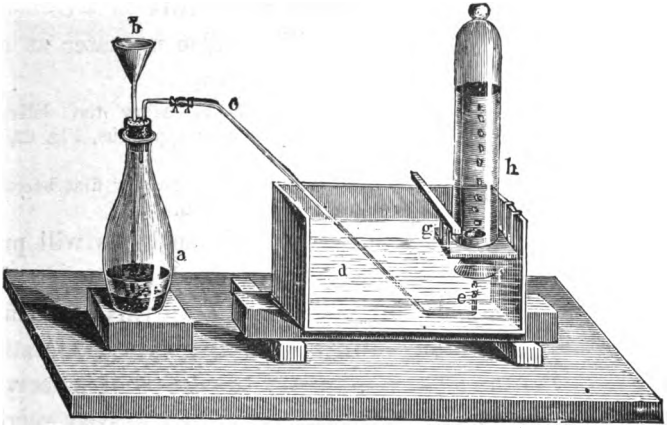

Fig. 63.—Preparation of carbonic acid gas.

We learn from this that carbonic acid gas will neither burn nor support combustion.

Experiment 102.—Place in a second jar of the gas a few strips of red and blue litmus papers which have been moistened with water.

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Observe that the blue litmus changes to a claret colour.

We learn from this that carbonic acid gas behaves with blue litmus as a **weak acid.**

Experiment 103.—Fit up an apparatus as shown in Fig. 49, arranging the jar of air with its mouth *upwards*. By means of a long delivery tube conduct carbonic acid gas into the jar.

Observe that as the gas displaces the air, the beam of the balance moves slowly downwards.

We learn from this that carbonic acid gas is **heavier than air.** Its specific gravity compared with air is 1·524.

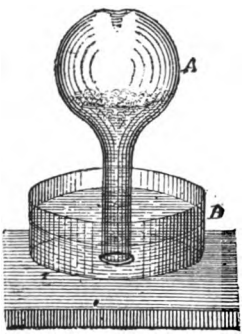

Fig. 64.—Solubility of carbonic acid gas in water.

Experiment 104.—Pour a little cold water into the flask which has been filled with carbonic acid gas, cork and shake well; then open the mouth of the flask under water which has been made blue with litmus.

Observe that the liquid rises in the flask, and that the water changes colour.

We learn from this that carbonic acid gas will dissolve in water at ordinary temperatures.

Experiment 105.—Allow the delivery tube from the generating apparatus, Fig. 62, to dip into lime-water.

Observe that the lime-water first becomes milky and then clears again.

We learn from this that carbonic acid gas will precipitate lime as **carbonate of lime,** and that when the gas is present in excess it will dissolve up this carbonate of lime and form **bicarbonate of lime;** this is the substance which causes the **temporary hardness of water.**

The amount of carbon dioxide present in a given weight of marble may be determined by means of the apparatus shown in Fig. 65. From one to two grams of powdered marble are introduced into the small flask. The flask is provided with a well-fitting indiarubber stopper, the

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delivery tube from which passes into a weighed U tube containing caustic soda, at the end of which is a small piece of tubing containing calcium chloride. When all the joints have been made secure, dilute hydrochloric acid is poured into the thistle funnel. Carbon dioxide is evolved and passes into the U tube containing the caustic soda, where it is absorbed; the tube at the end containing the calcium chloride absorbs any moisture. It is best, also, in order to obtain exact results, to place another U tube containing calcic chloride between the flask and the caustic soda tube, in order that the liberated gas may be thoroughly dried. The increase in weight of the U tube represents the carbon dioxide which was contained in the marble. During the evolution of the gas fresh acid must be added to the marble from time to time to ensure the decomposition of the whole of the marble and the liberation therefore of all the carbon dioxide. The results of many experiments conducted with great care, but upon this plan, have shown that in all cases 100 *parts by weight of pure carbonate of lime contain* 44 *parts by weight of carbon dioxide.*

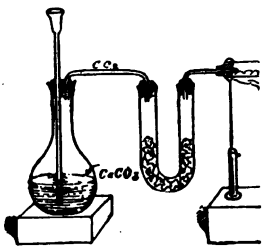

Fig. 65.—Percentage of carbon dioxide in marble.

The **composition of carbonic acid gas** has been determined by burning a known weight of pure carbon, diamond, in pure oxygen, and absorbing the whole of the carbonic acid gas produced. The apparatus is shown in Fig. 66. Where A is a gasholder containing oxygen, and B, C, tubes containing caustic potash, by which the oxygen is dried and purified, D E is the piece of combustion

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tubing containing a weighed quantity of pure carbon; this is placed in a small furnace, and the carbonic acid gas is caught in the weighed tubes F, G, and H, of which F and H contain solid caustic potash, and G a strong solution of the same body. The increase in weight of the tubes F, G, and H, will give the amount of carbonic acid gas produced, and the decrease in weight of D, E, the weight of carbon burnt. From this we may calculate at once the oxygen which has united with the carbon. The results of many experiments show that 100 *parts by weight of carbonic acid gas contain* 27·27 *parts by weight of carbon,*

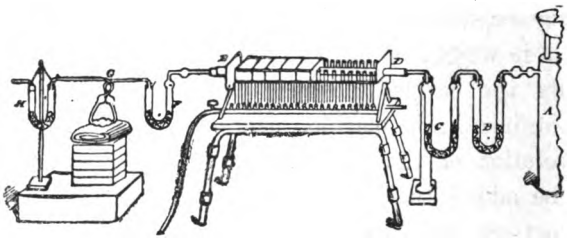

Fig. 66.—Apparatus for determining the composition of carbon dioxide.

and 72·73 *parts by weight of oxygen.* Since another oxide of carbon is known which contains just half this amount of oxygen in proportion to carbon, it is called **carbon monoxide**; and carbonic acid gas, because it contains twice as much oxygen in proportion to carbon as this lower oxide, is known as **carbon dioxide.**

CARBON MONOXIDE is produced from carbon dioxide by passing it over carbon heated to redness. The pale blue flames which play over the top of a coke, charcoal, or anthracite fire are due to the presence and combustion of this gas. In the lower part of the charcoal fire the air

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oxidises the carbon, as we saw in experiment 100; the carbon dioxide thus produced passes over the red-hot charcoal, and is there converted into carbon monoxide. This body escapes from the top of the fire, and, coming in contact with more oxygen, burns with a characteristic flame. This action of a charcoal fire may be well illustrated by means of the apparatus shown in Fig. 67. Carbon dioxide

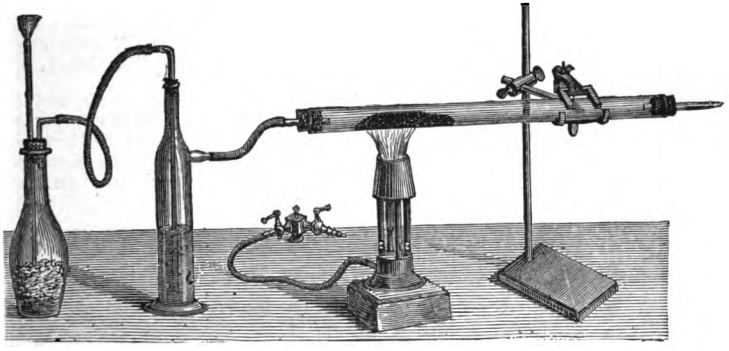

Fig. 67.—Preparation of carbon monoxide from carbon dioxide.

is generated in the flask on the left; it is then purified by passing it through sulphuric acid, and allowed to flow over the red-hot carbon in the combustion tube. Combustible carbon monoxide escapes from the delivery tube.

The **properties of** carbon monoxide may be demonstrated by collecting a few small jars of the gas as prepared by the method above described. On examination it proves to be colourless, with a faint, oppressive odour, and lighter than air. It does not support combustion, but burns with a pale blue flame, *combining with oxygen and forming, during combustion, carbon dioxide.*

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SUMMARY TO CHAPTER VIII.

CARBON.

- **Occurrence in Nature** — In the inorganic kingdom it occurs combined as carbonates, and in the air as carbon dioxide. Free, it exists in nature as *diamond*, *graphite*, and *jet*, and, in an impure form, in all the varieties of coal. It is present in all organic things.
- **Properties** — Solid body, which burns when heated in air or oxygen to form *carbon monoxide* or *carbon dioxide*. It is not acted upon by *acids* or *alkalies*; decoloriser, deodorant, and disinfectant.
- **Preparation** — From coal, wood, and bone by destructive distillation, as coke, wood charcoal, and bone charcoal; and from resins, as lamp black.

CARBONIC ACID GAS, OR CARBON DIOXIDE, OR CARBONIC ANHYDRIDE.

- **Occurrence in Nature** — It exists in small quantities free in air, and is always produced by the combustion of organic substances. Combined with bases it is largely present in the mineral kingdom in the form of carbonates. The gas is also dissolved in the waters of many springs.
- **Properties** — A colourless gas with slight taste, will not burn and is a non-supporter of combustion, soluble in water, heavier than air.
- **Tests** — It turns litmus claret colour, and yields a white precipitate of calcic carbonate with lime-water. The carbonates effervesce on the addition of hydrochloric acid, and carbon dioxide is liberated.
- **Preparation** — By pouring hydrochloric acid on chalk, marble, limestone, and oyster shells, and by direct combination with oxygen when carbon is burnt.

CARBON MONOXIDE, OR CARBONIC OXIDE.

- **Occurrence in Nature** — The gas does not exist in the free state in nature, but is produced during the incomplete oxidation of carbon, as in a charcoal fire.
- **Properties** — A colourless, almost inodorous, tasteless gas, which burns in air, forming carbon dioxide. It is extremely poisonous.
- **Tests** — No action on lime-water, but burns in air to produce carbonic acid gas, which turns lime-water milky.
- **Preparation** — By passing a current of carbon dioxide over red-hot charcoal.

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QUESTIONS ON CHAPTER VIII.

1. How does carbon occur in nature?

2. Describe the forms of carbon. How would you prove that each of these substances consist of the same element?

3. How would you prove that sugar, gum, camphor, white of egg, and turpentine all contain carbon?

4. What changes take place in the transformation of wood into coal?

5. What are some of the common uses of carbon?

6. What body is produced by the combustion of carbon in air?

7. You are required to prepare carbon dioxide; give a description of one process, and a sketch of the apparatus.

8. What weight of carbon and of oxygen is there in 100 parts by weight of each of the oxides of carbon?

9. Quicklime is slaked with water, diffused through more water, and filtered. When carbonic acid gas is passed through the filtered liquid it becomes very turbid; what is the composition of the substance which causes the turbidity?

10. Carbonic acid gas (carbon dioxide) is passed into (*a*) caustic potash, (*b*) lime-water. State what happens in each case, and give the names of the substances produced.

11. Both water and carbon dioxide are formed when coal gas burns; describe, and show by sketch, how you would prove this to be the case.

12. Describe fully how you would prepare carbon monoxide from carbon dioxide.

13. How is calcium bicarbonate produced?

14. Air passes through a bright coke fire; describe the chemical changes which will occur and the properties of any compounds that may be formed.

15. Write in two parallel columns the respective properties of carbon monoxide and dioxide. How can the monoxide be converted into the dioxide, and the dioxide into monoxide?

16. Lime-water is shaken up in a jar filled with carbonic oxide; what occurs? The carbonic oxide is then inflamed and the lime-water once more shaken up with the contents of the jar. What now takes place?

17. Sketch the apparatus you would employ for the preparation of carbon monoxide.

18. How may carbon monoxide be converted into carbon dioxide?

19. How would you prepare carbonic anhydride (carbonic acid)? Give a sketch of the apparatus and describe the properties of this gas.

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Preparation for Ninth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of blende, iron pyrites, galena, roll sulphur, flowers of sulphur, native sulphur, retort, hard-glass tubing, test-tubes, flask, crucible, sand-bath and tripod flask, copper turnings, strong sulphuric acid, gas jars and plates, litmus, apparatus as in Figs. 74, 75, 76, platinum chloride, asbestos, ferrous sulphide, ether, beakers, balance, corks, cork-borers, and delivery tubing.

EXPERIMENTS TO BE PERFORMED.

"Show roll and flowers of sulphur and specimen of native sulphur, also iron pyrites and other native sulphides.

Powder iron pyrites and heat the powder in a test-tube held horizontally over a lamp to show the sulphur obtained from the pyrites.

Show the melting of sulphur by heating flowers of sulphur in a small flask.

Heat sulphur on a piece of tin plate till it catches fire, show the colour of the flame and observe the smell of burning sulphur.

Prepare sulphur dioxide by heating copper turnings in strong sulphuric acid, and show that it extinguishes flame, is very soluble in water, and that the water dissolving it becomes very acid, turning blue litmus red.

Bubble air through a strong solution of sulphur dioxide, and then over platinised asbestos; demonstrate that when the platinised asbestos is hot dense white fumes are formed, which consist of sulphur trioxide

Pour some sulphuric acid into 20 or 30 times its volume of water and prove its acid taste, its action on litmus, and its power of causing effervescence if dropped on sodium carbonate.

Show that sulphuric acid is a colourless liquid, that bulk for bulk it is more than $1\frac{3}{4}$ times the weight of water.

Show by a thermometer or by immersing a test-tube with spirit in it that a large amount of heat is evolved when this acid is poured into water.

Pour some sulphuric acid on sugar or shake it up with oil to show its action on organic bodies."—*Science Department's Syllabus.*

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CHAPTER IX.

SULPHUR OR BRIMSTONE.

Where found—Its properties—Is also found chemically combined with many Metals, so not recognisable by the eye—Sulphur heated in the Air melts; more strongly heated it burns, then the Sulphur disappears; the strong smell produced belongs to a new body formed by the burning; this new body is a compound of Sulphur and Oxygen—Gaseous properties of the new body, its effect on Blue Litmus Paper, which Oxygen and Sulphur have not—Its composition is 50·00 *parts of Sulphur and* 50·00 *parts of Oxygen, and it is called Sulphur Dioxide—Water dissolves nearly fifty times its volume of this Gas, and then turns Blue Litmus strongly Red and has an acid taste—The combination of the Gas with Water to form Sulphurous Acid—Another compound of Sulphur and Oxygen can be made, in which the same weight of Sulphur is combined with more Oxygen—One hundred parts contain* 40 *of Sulphur and* 60 *of Oxygen, and it is called Sulphur Trioxide—Sulphur Trioxide has properties differing from the Dioxide—If the Dioxide and Oxygen be mixed they do not combine, but if they are passed together over hot platinised asbestos dense white fumes are formed which are the Trioxide—Combination of the Trioxide with Water to form Sulphuric Acid (Oil of Vitriol).*

SULPHUR is the solid, bright yellow body which is commonly known as **brimstone.** It is found native or free in the neighbourhood of volcanoes; large quantities are obtained from the volcanic areas of Italy, Sicily, Iceland, and South America. When pure it is a brittle yellow substance which will burn in the air with a characteristic blue flame. Sulphur also exists in large quantities in the earth, combined with many metals; such bodies are known as **sulphides,** and since they consist of sulphur in chemical union with other bodies, it is easily

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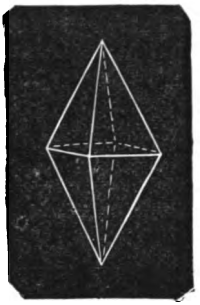
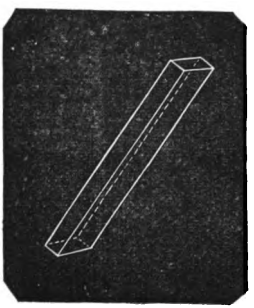
understood why the sulphur is in such cases not recognisable by the eye. **Sulphide of hydrogen** is a gas which is given off during the decomposition of organic bodies, and which is often liberated from volcanoes. Combined with oxygen and a metal, sulphur is found as **sulphates,** and united with a smaller proportion of oxygen and a metal, it exists in **sulphites.** Many of the substances found in living things contain sulphur, and it is also present in a great number of minerals.



A
Fig. 68.—Octahedral crystal.

Sulphur and *iron* form **iron pyrites.**
„ „ *zinc* „ **blende.**
„ „ *lead* „ **galena.**
„ „ *mercury* „ **cinnabar.**
„ „ *copper* „ **copper glance.**
„ „ *silver* „ **silver glance.**

Sulphur united with *oxygen* and *calcium* forms **gypsum.**
„ „ „ „ *sodium* „ **Glauber salts.**
„ „ „ „ *barium* „ **heavy spar.**
„ „ „ „ *magnesium* „ **Epsom salts.**



B
Fig. 68.—Prismatic crystal.

Native sulphur occurs as more or less perfect **octahedral crystals,** Fig. 68A. This variety is *soluble in carbon disulphide.*

When ordinary sulphur is melted and allowed slowly to cool, it forms long needle-shaped **prismatic crystals.** These crystals are at first slightly elastic, but rapidly become opaque and brittle. This variety is *insoluble in carbon disulphide* (see Fig. 68B).

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Another form of sulphur is prepared by pouring melted sulphur, which has been heated to 24° C., into cold water. This variety is a soft **plastic body** which is only *slightly soluble in carbon disulphide.*

Amorphous sulphur is prepared by partially dissolving the plastic sulphur in carbon disulphide, when an insoluble powder non-crystalline in character is left.

Preparation.—Free sulphur is usually found in nature associated with earthy substances, such as clay, gypsum, etc. From these it is isolated by roasting the ore in large heaps. The sulphur ore is arranged in stacks through which shafts run, and the whole is covered in with sods of earth (Fig. 69). The sulphur is then ignited below, the oxidation is limited because of the small amount of air admitted, and the heat produced serves to melt out the sulphur from the remaining mass. This **crude sulphur** is further purified by **sublimation.**

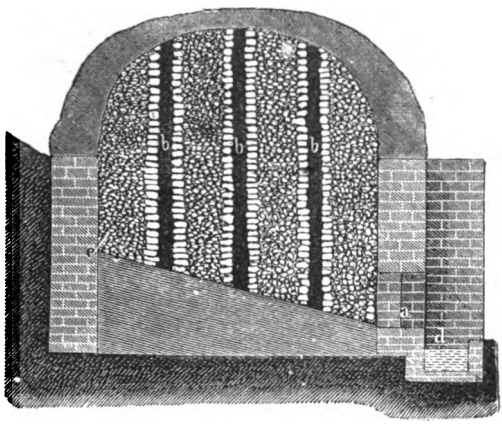

Fig. 69.—Preparation of sulphur.

The method adopted in extracting sulphur from its ore may be imitated on a small scale by the following :—

Experiment 106.—Place in a test-tube or retort a little *sulphur ore*, or if this cannot be obtained use powdered sulphur mixed with gypsum and clay. Heat the test-tube or retort.

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Observe that, on the application of heat, a vapour is produced which condenses on the sides of the retort.

We learn from this that sulphur may be separated from earthy bodies such as clay by sublimation.

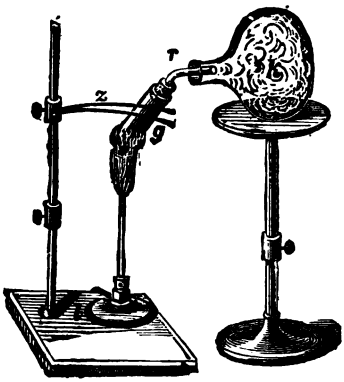

Fig. 70.—Sublimation of sulphur.

Sulphur may also be obtained from certain sulphides which are rich in this body by the action of heat.

Experiment 107.—Introduce into a piece of combustion tubing closed at one end a small quantity of finely-ground *iron pyrites*. Strongly heat the tube.

Observe that the sulphur sublimes in the colder part of the tube.

We learn from this that sulphur may be obtained from iron pyrites by the action of heat.

Crude sulphur obtained by roasting the ore is purified

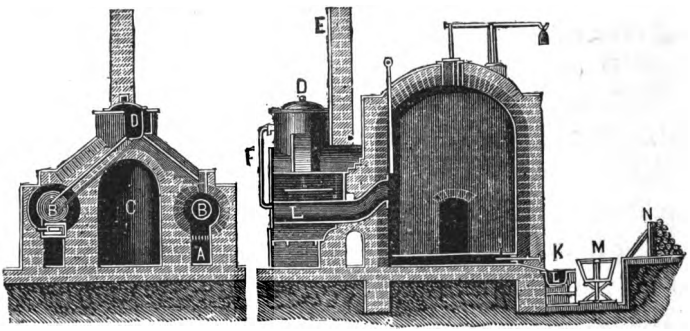

Fig. 71.—A sulphur chamber.

for commercial purposes by distillation. It is heated in large iron retorts, Fig. 71, L; the vapour condenses on

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the walls of a brick chamber as a fine yellow powder, which finds its way into the market as **flowers of sulphur.** On the floor of the condensing chamber some liquid sulphur collects; this is drawn off and run into cylindrical moulds, and on cooling forms the sticks known as **roll sulphur.**

When sulphur is heated it undergoes a somewhat remarkable set of changes, but finally it boils, and if exposed to air takes fire, burning with a blue flame and yielding a heavy poisonous gas which possesses a strong suffocating odour.

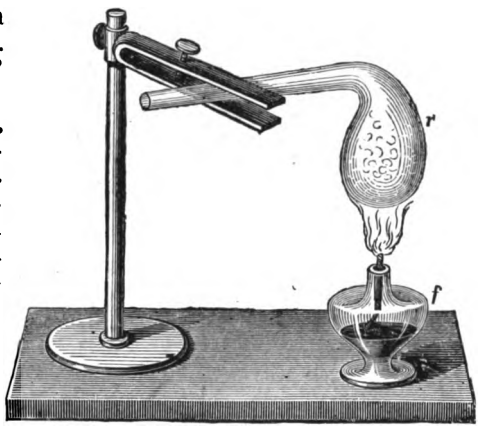

Fig. 72.—Action of heat upon sulphur.

Experiment 108.—Gently heat some *sulphur* in a retort, Fig. 72.

Observe that the sulphur melts and forms an **amber-coloured mobile liquid,** which flows easily as the retort is turned from side to side. As the application of heat is continued the liquid darkens and thickens, until the mass becomes so thick and **viscid** that the retort may be inverted without causing the sulphur to run out. On further heating, the body becomes a thin, dark-brown liquid, and finally **boils,** giving off **dark-red vapours** which condense to a yellow powder in the neck of the retort. The vapour takes fire at the mouth of the retort.

We learn from this that sulphur is peculiar in its behaviour on the application of heat.

Carefully conducted experiments have shown that sulphur first melts at about 115° C.; at 138° C. it begins to turn brown, and becomes a nearly solid substance at 230° C.; it assumes the brown liquid condition at 260° C., and boils at 440° C.

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Experiment 109.—Place a piece of sulphur on an iron plate, and heat the plate by means of a Bunsen flame.

Observe that the sulphur melts and passes through all the changes above described, and when more strongly heated it *takes fire in the air*, burning with a blue flame, and giving off a strong-smelling gas, which will at once redden moistened blue litmus paper.

We learn from this that sulphur is **inflammable** when heated in air, and on burning yields a gas. That the sulphur disappears and a new body is formed by its combination with oxygen of air. This new substance is unlike oxygen or sulphur, for it has an acid reaction, as is shown by its action upon blue litmus. This gas is called **sulphur dioxide**, and it has been shown by exact experiments that *it contains in* 100 *parts by weight*, 50 *parts of sulphur and* 50 *parts of oxygen.*

SULPHUR DIOXIDE is a colourless, pungent, suffocating gas, which may be formed by the combustion of sulphur in air or oxygen, or by roasting metallic sulphides in air; but when so prepared it is always impure, for it is mixed with nitrogen. The gas is not inflammable, and quickly extinguishes the flame of a burning body. It is very soluble in cold water, and must therefore be collected by the displacement of air, or over mercury. *Cold water will absorb nearly* 50 *times its volume of the gas.* A solution of the gas in water is known as **sulphurous acid.** The gas is a powerful **bleaching agent**; it is useful for removing the stains of fruit or red wine from linen, and it may be employed with good results in most cases where chlorine acts injuriously upon the material. Commercially, it is employed in bleaching straw and woollen goods. Its bleaching properties depend upon the

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fact that it has a powerful affinity for oxygen. In the presence of water it fixes the oxygen and liberates hydrogen; this liberated hydrogen decomposes the colouring matter, forming a colourless compound. In consequence of this

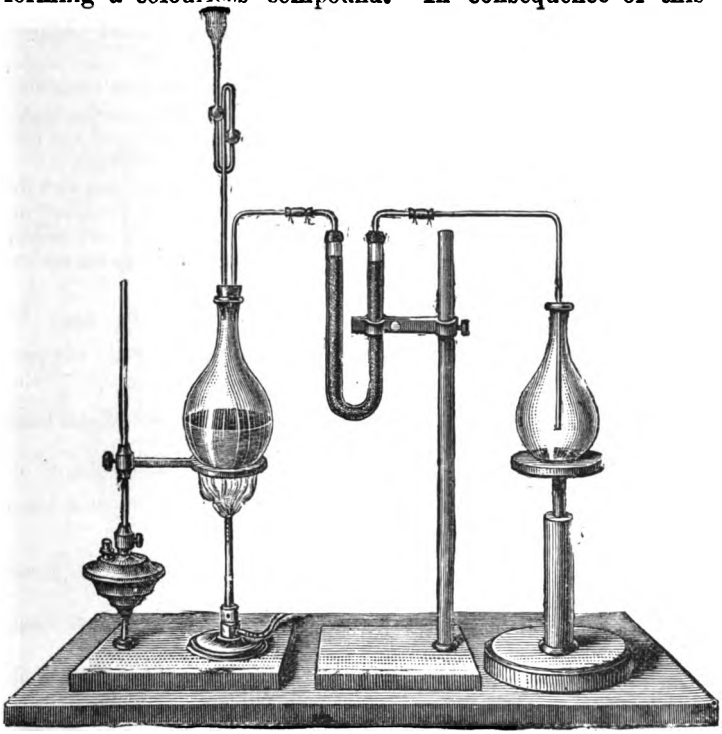

Fig. 78.—Preparation of sulphur dioxide.

great chemical affinity for oxygen, sulphur dioxide acts as a **disinfectant** and an **antiseptic**. The gas may be liquefied by passing it into a tube surrounded by a freezing mixture.

Sulphur dioxide is met with in the gases evolved by volcanoes, and it is usually present in small quantities in

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the air of towns and manufacturing districts; in the latter case, it mainly owes its origin to the oxidation of the sulphur of the iron pyrites present in coal.

Preparation.—A supply of the gas may be prepared with ease by the action of sulphuric acid on heated copper.

Experiment 110.—Introduce into a flask, Fig. 73, some copper turnings, connect the delivery tube of the flask with a U tube containing pumice-stone saturated with sulphuric acid. Now pour down the thistle tube sufficient strong sulphuric acid *to cover* the metallic copper, and heat gently. Collect the gas which comes off by downward displacement.

Observe that effervescence occurs, and that an invisible gas, with the characteristic *odour of burning sulphur*, comes off. Attention should be paid to the apparatus during heating, for the gas is liberated very rapidly and the liquid froths up. If, therefore, the effervescence is very violent the Bunsen flame should be removed.

We learn from this that sulphur dioxide may be obtained by the deoxidising action of the metal, copper, upon sulphuric acid.

Experiment 111.—To the mouth of a jar of *sulphur dioxide* bring a burning taper, and then plunge the taper into the jar.

Observe that the gas does not burn, and the flame is extinguished.

We learn from this that sulphur dioxide does not burn nor does it support combustion.

Experiment 112.—Open a jar of sulphur dioxide with its mouth immersed in a vessel of water coloured with blue litmus.

Observe that the liquid rises in the jar, and that the solution of litmus is turned quite red.

We learn from this that sulphur dioxide is soluble in water, and that it then forms an acid body.

Water will absorb about fifty volumes of sulphur dioxide, and the resulting substance is **sulphurous acid.** A solution of this body may be prepared by passing the gas into water by means of the apparatus shown in Fig. 74. Sulphurous acid decomposes at a low heat, or on agitation into water and sulphur dioxide.

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SULPHUR TRIOXIDE.—Another compound can be prepared of sulphur and oxygen, in which the same weight of sulphur is combined with more oxygen. This body is produced by the chemical union of sulphur dioxide and oxygen, but these substances will not directly combine under ordinary circumstances; other means have therefore to be resorted to in order to produce sulphur trioxide.

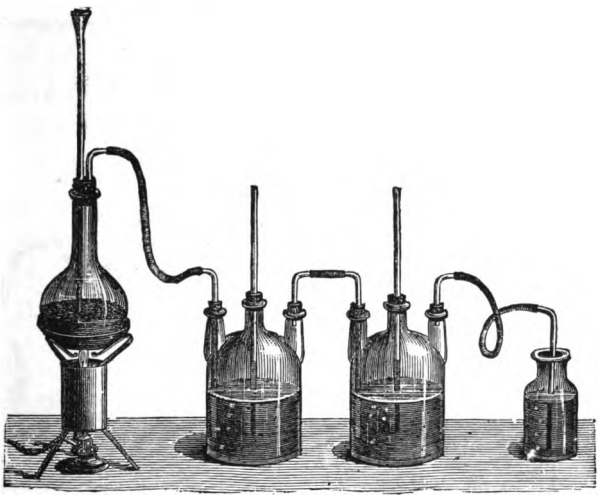

Fig. 74.—Preparation of sulphurous acid.

The method which may here be adopted is to pass sulphur dioxide, mixed with air or oxygen, over heated platinized asbestos. The latter body is prepared by soaking asbestos in a strong solution of platinum chloride, and then igniting the resulting substance. By this means the platinum chloride is decomposed, and finely divided platinum is deposited upon the asbestos. The sulphur dioxide may be conveniently obtained by passing a current of air

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through a strong solution of sulphurous acid. The sulphur dioxide, when mixed with air and passed over the heated asbestos, takes up oxygen, and thus a new body is formed which yields heavy white fumes which condense to produce needle-shaped crystals of sulphur trioxide.

Experiment 113.—Fit up a piece of apparatus as shown in Fig. 75, so that air is made to pass from the aspirator, on the left, through a strong solution of sulphur dioxide in water, the delivery tube from the bottle conducts the mixture of air and sulphur dioxide into a retort containing

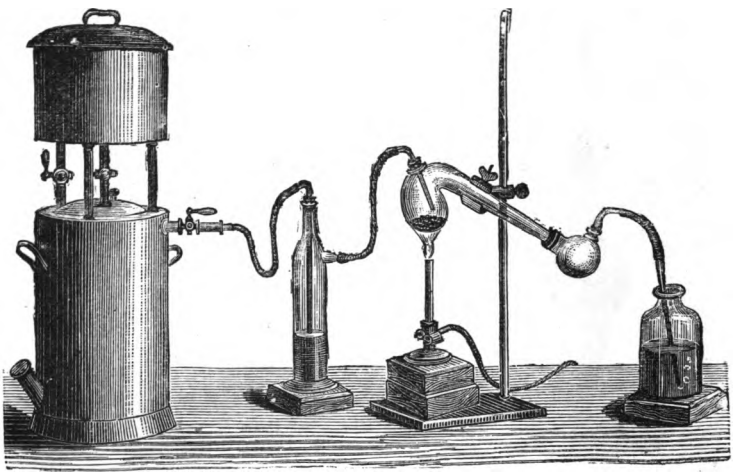

Fig. 75.—Preparation of sulphur trioxide.

heated platinized asbestos. On the right the retort is fixed into a glass bulb, from which a delivery tube passes into water.

Observe that on passing air from the aspirator, through the solution of sulphurous acid, over the heated, finely-divided platinum, which has been deposited upon the asbestos, heavy white fumes are produced, some of which condense in the bulb and others pass into the water. The water becomes strongly acid in reaction.

We learn from this that sulphur dioxide and oxygen may be made to combine with oxygen, when the two bodies are passed over finely divided platinum.

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Dry solid sulphur trioxide combines most readily with water, forming **sulphuric acid.** For this reason, the solid body liquefies on exposure to air, from which it absorbs water. If a few pieces of sulphur trioxide are thrown into water a loud hissing noise is produced, like that made by plunging red-hot iron into water. The **oil of vitriol** of commerce is strong sulphuric acid. It is an oily-looking, inodorous, colourless liquid. It is one of the most active of the acids, and, from its powerful affinity for moisture, chars most organic substances with which it is brought in contact. So great is its affinity for water in any form, that if a small quantity of the concentrated acid is exposed to the air of a room, it will frequently double its weight in a few days. For this reason, strong sulphuric acid is employed as a **desiccating agent** to dry air and other gases. In mixing sulphuric acid with cold water, care should always be taken that the strong acid is *added to the water*, gradually, with constant stirring.

Experiment 114.—Place a little *strong sulphuric acid* in a test-tube, and then add a piece of loaf sugar.

Observe the heavy, oily nature of the acid, and that it chars sugar.

We learn from this that oil of vitriol is a colourless, inodorous, heavy liquid, oily in appearance, and that it chars the organic body, sugar.

Experiment 115.—Measure into a beaker about 25 cc. of oil of vitriol, and carefully note the weight; then into another beaker introduce the same volume of distilled water, and weigh. Deduct in each case the weight of the empty glass.

Observe that the sulphuric acid is bulk for bulk heavier than water.

We learn from this that sulphuric acid is, volume for volume, about $1\frac{3}{4}$ times heavier than water.

Experiment 116.—Pour some *dilute sulphuric acid*, as obtained in

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experiment 113, Fig. 75, upon some dry solid soda carbonate in a test-glass, and another portion add to some blue litmus solution.

Observe that the soda carbonate on the addition of the acid violently effervesces, and that the litmus is reddened.

We learn from this that the body has acid properties.

Experiment 117.—Fit up a test-tube with a sound cork and delivery tube drawn out to a jet, and place a small quantity of *ether* in the tube. Now prepare some dilute sulphuric acid by adding strong oil of vitriol to cold water drop by drop with constant stirring; the test-tube containing the ether may be employed as a stirring rod.

Observe that the solution becomes very hot and steam is given off, and the ether is converted into vapour, which will ignite at the jet of the tube on the application of a burning match.

We learn from this that sulphuric acid has a great affinity for water, and that much heat is evolved on the preparation of a solution of the body.

SULPHURETTED HYDROGEN is a compound of sulphur and hydrogen, which is of very great importance in chemical work. This body is a colourless gas with a strong offensive odour, somewhat like that of rotten eggs. It burns in the air with a blue flame, but will not support combustion or life. The gas is soluble in cold water; it should therefore be collected over warm water, in which body it is only very slightly soluble. The solution of sulphuretted hydrogen has a peculiar odour, and since it has a slightly acid reaction it is known as **hydrosulphuric acid.** In the laboratory of the chemist it is one of the most important reagents, for by means of it many metals are converted into **insoluble sulphides**; and as such may be precipitated from their solutions by the action of this gas. Sulphuretted hydrogen occurs in minute quantities in the air of towns and manufacturing districts; it has also been detected in the gases evolved from volcanoes. During

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the combustion of coal and decomposition of animal and vegetable substances the gas is produced in small quantities. The mineral sulphides—compounds of a metal with sulphur—which are so very abundant in nature, may be looked upon as substances in which the hydrogen of sulphuretted hydrogen has been displaced by a metal.

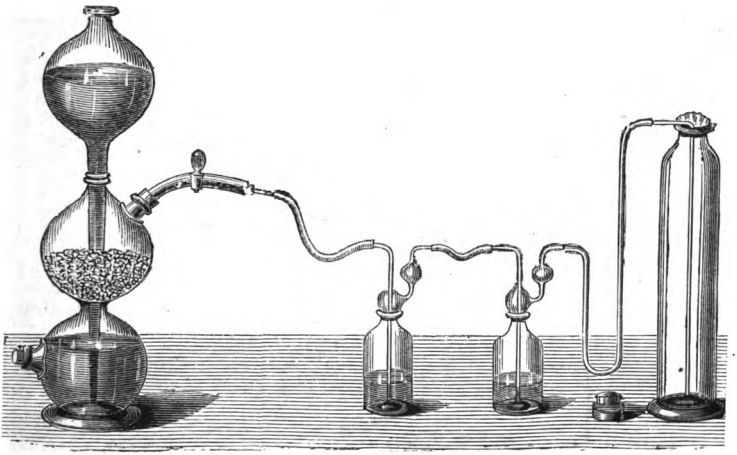

Fig. 76.—Preparation of sulphuretted hydrogen.

This gas may be most conveniently prepared by the action of sulphuric acid upon sulphide of iron.

Experiment 118.—Fit up an apparatus as shown in Fig. 76. Place in the middle bulb of the apparatus on the left some broken pieces of *sulphide of iron*, and pour into the upper bulb a dilute solution of *sulphuric acid*. The gas may be washed by means of water placed in the two bottles.

Observe that a gas escapes with a strong odour like that of rotten eggs.

We learn from this that sulphuretted hydrogen is a gas which may be produced by the action of sulphuric acid upon sulphide of iron; that it is invisible, and has a strong offensive odour.

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SUMMARY TO CHAPTER IX.

SULPHUR

- **Occurrence in Nature**
  - It is present, combined with other elements, in many animal and vegetable substances.
  - In the earth it is found free, also largely combined with metals, as *sulphides*, *sulphates*.
  - Traces of *sulphites* exist in some natural waters.
- **Properties**
  - Sulphur is a yellow, brittle, inflammable solid. It is more or less soluble in turpentine, carbon disulphide, sodic hydrate, and alcohol; but it is insoluble in water.
  - Several allotropic forms of sulphur are known, as the *rhombic*, *prismatic*, *plastic*, *amorphous*.
  - All the varieties burn in air to form sulphur dioxide. Sulphur when heated will combine with most metals and with many non-metals.
- **Preparation**
  - Sulphur is obtained chiefly from its ores by sublimation.
  - **Brimstone or commercial roll sulphur** is prepared by melting sulphur, after which it is cast into sticks.
  - **Flowers of sulphur** produced when the vapour of sulphur is quickly cooled below its melting point.

SULPHUR DIOXIDE

- **Occurrence in Nature**
  - Sulphur dioxide is present in volcanic gases, and traces exist in the air.
- **Properties**
  - Colourless gas, with a strong pungent odour, will not burn nor support combustion. Soluble in water, and the solution forms *sulphurous acid*.
- **Preparation**
  - (1) By burning sulphur in air or oxygen.
  - (2) By heating sulphuric acid and copper.

SULPHUR TRIOXIDE

- **Occurrence in Nature**
  - Does not occur free, but in the combined form as sulphates.
- **Properties**
  - A white crystalline solid body, which combines with water to form *sulphuric acid*.

SULPHURIC ACID

- **Occurrence in Nature**
  - Does not occur free, but combined with metals, and in solution as soluble sulphates in some natural waters.
- **Properties**
  - A heavy non-fuming, oily liquid, it combines with water in all proportions, and evolves great heat during the combination.

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QUESTIONS ON CHAPTER IX.

1. What sulphur compounds occur in nature? Describe the chief properties of sulphur, and state how it may be converted into its various allotropic modifications.

2. Describe the changes produced in brimstone by the continued application of heat.

3. How does native sulphur occur, and how is it prepared from its ore?

4. How would you prepare sulphur from iron pyrites?

5. Write down the names of ten common minerals which contain sulphur.

6. What are the chief chemical and physical properties of sulphur?

7. What is the difference between a sulph*ate* and a sulph*ite*?

8. Sketch a sulphur chamber. Name the parts and explain the use of each.

9. If I burn a piece of sulphur in a bottle filled with air, and another piece in a bottle filled with oxygen, what shall I find in each of the bottles after the combustion is finished?

10. What happens when sulphuric acid and copper turnings are heated together? Sketch the apparatus used for collecting the product, and describe its properties.

11. Sulphur is burnt in air, and the product passed into water. State what may be the products and how they may be detected.

12. How is sulphur dioxide met with in nature?

13. How could you convert sulphur dioxide into sulphuric acid, and sulphuric acid into sulphur dioxide?

14. How would you distinguish a solution of sulphurous acid from one of sulphuric acid?

15. Sketch and describe the apparatus used for preparing a small quantity of sulphuric acid in the laboratory.

16. What are the general properties of sulphuric acid?

17. How would you prepare some dilute sulphuric acid from oil of vitriol?

18. How would you prepare sulphuretted hydrogen?

19. What are the chief properties of sulphuretted hydrogen?

20. Sketch the apparatus you would require for the preparation of sulphurous acid.

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Preparation for Tenth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of rock salt and common salt, apparatus as shown in fig. 77, gas jars and plates, manganese dioxide, salt, sulphuric acid, hydrochloric acid, Dutch metal, turpentine, cotton wool, phosphorus, litmus, coloured calico, sodium, hydrogen apparatus, bleaching powder, three large basins, deflagrating spoons.

EXPERIMENTS TO BE PERFORMED.

"Prepare chlorine from (1) mixture of common salt, black oxide of manganese, and sulphuric acid;

(2) from mixture of black oxide of manganese and hydrochloric acid. Collect gas by downward displacement.

Draw attention to its colour, and show that phosphorus spontaneously inflames in the gas to form a chemical compound of phosphorus and chlorine.

Show that oil of turpentine ignites spontaneously in chlorine.

Show that sodium when strongly heated and immersed in a jar of chlorine burns and forms common salt.

Show bleaching action of chlorine by dipping moistened Turkey red rag in bottle filled with gas.

Show similar action with solution of bleaching powder and acid.

Show that chlorine is soluble in water and that the solution has characteristic smell and colour of the gas."—*Science Department's Syllabus.*

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CHAPTER X.

CHLORINE.

The Gas obtained by the action of Hydrochloric Acid on the Black Oxide of Manganese (so called on account of its colour)—Its characteristic smell —Is 2½ times as heavy as Air and 35½ times as heavy as Hydrogen— Soluble in Water—Many substances take fire in Chlorine Gas, e.g., Phosphorus, and form Chlorides—Ignition of Oil of Turpentine in Chlorine with separation of Carbon and formation of Hydrochloric Acid—Bleaching Power of Chlorine—Bleaching Powder.

CHLORINE is a word derived from *chloros*, meaning green. It is the name given to a greenish-yellow gas which may be got out of common salt, and which may also be extracted from a great many other bodies known as **chlorides.** This gas has a very strong suffocating odour; it acts as a violent irritant even when breathed in small quantities, and produces prolonged fits of violent coughing. Whilst it will support the combustion of many bodies it will not burn in air, although metals will burn in it; in fact it possesses great chemical affinity for both metals and non-metals, and combining with them forms chlorides. In the presence of water chlorine is a powerful bleaching agent, and it is the important and active constituent of **bleaching powder.** The gas is somewhat soluble in cold water, for at 15° C. even, water will absorb 2·37 volumes of chlorine; but as it is only slightly soluble in hot water and is heavier than air, it should be collected either over warm

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water or by the displacement of air. Chlorine is 2½ times heavier than air, and 35½ times as heavy as hydrogen. It is not met with free in nature, but is found in combination with metals, chiefly with sodium as common salt. From this body the chlorine may be prepared by first withdrawing the other substance by the use of suitable reagents.

Experiment 119.—Prepare a mixture of *black oxide of manganese* and *fused salt*, and introduce it into a flask, as shown in Fig. 77. Add sulphuric acid, and heat the mixture.

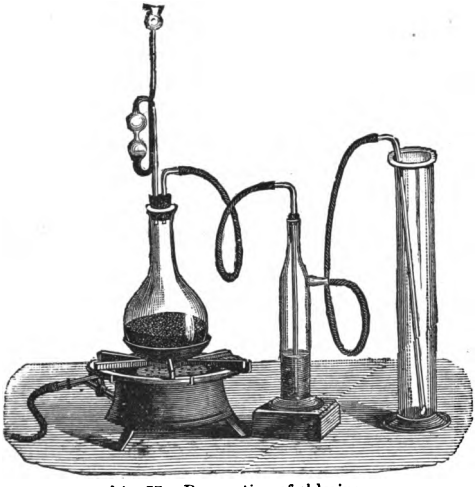

Fig. 77.—Preparation of chlorine.

Observe that effervescence occurs, and a greenish gas escapes which possesses a strong suffocating odour.

We learn from this that chlorine may be prepared from common salt by the action of manganese dioxide and sulphuric acid. As this mixture is somewhat liable to froth up and boil over, it is better to prepare the gas by the action of manganese dioxide upon hydrochloric acid.

Experiment 120.—Into a 10-oz. flask, fitted up as shown in Fig. 77, introduce *strong hydrochloric acid* until it is about a quarter full; then add of *black oxide of manganese* about one-sixth part by weight. Apply gentle heat, and collect six jars of the gas which has been dried by the action of the sulphuric acid contained in the wash-bottle.

Observe that the reaction commences when the mixture is cold, and chlorine gas is liberated, which may be collected by downward displacement,

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as shown. The rate of evolution of the gas increases on warming the flask. The gas is heavier than air, for it is with ease collected by displacement of air. It is greenish yellow in colour when examined by daylight, and has a peculiar irritating odour. If prepared at night a jar should be examined by the light of a piece of burning magnesium ribbon.

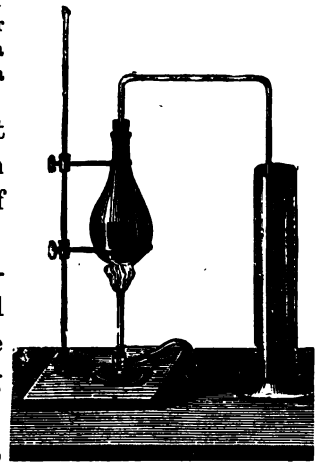

Fig. 78.—Preparation of chlorine water.

We learn from this that chlorine may be prepared from hydrochloric acid by the action of manganese dioxide.

Chlorine gas, as above explained, cannot well be collected over cold water as it is soluble in that body; but by passing a current of chlorine through cold water, as shown in Fig. 78, a solution of the gas is formed which has a pale yellowish tint and the odour of chlorine. This solution is known as **chlorine-water.**

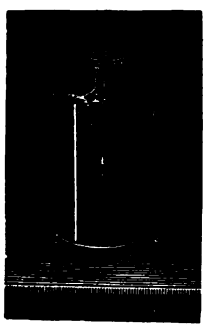

Fig. 79.—Phosphorus burns in chlorine.

Experiment 121.—Cut a small piece of *phosphorus* and dry it on the folds of a filter paper, place it in a deflagrating spoon and immerse it into a jar of chlorine.

Observe that the solid phosphorus melts, and afterwards takes fire spontaneously, and burns with a small pale-blue flame; dense fumes are liberated, and a yellow deposit of phosphorus pentachloride is formed on the sides of the jar.

We learn from this that the non-metal, phosphorus, has a very strong affinity for chlorine, for it takes fire spontaneously in the gas.

Experiment 122.—Introduce a *burning taper* into a jar of chlorine.

Observe that the flame of the taper becomes lurid red, and dense white fumes are produced and clouds of black smoke are liberated.

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We learn from this that a taper will burn in chlorine, and we shall see that the altered nature of the flame is due to the fact that the chlorine combines with the hydrogen only of the fat to form dense, white, fuming, hydrochloric acid gas, and that it is the carbon of the fat which is liberated to form the clouds of smoke.

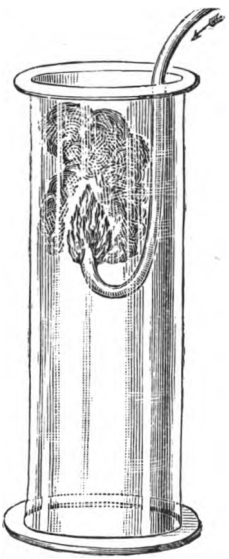

Fig. 80.—Hydrogen burns in chlorine.

Experiment 123.—Fit up an apparatus, as shown in Fig. 80, so that a delivery tube from the hydrogen apparatus, ending in a small jet, may be introduced into a jar of chlorine. After testing the escaping hydrogen by filling a test-tube and observing that the gas burns with a quiet flame at the mouth, ignite the gas and introduce the flame into the jar of chlorine.

Observe that the gas continues to burn, but that clouds of fumes are formed and the yellow chlorine disappears.

We learn from this that chlorine will combine directly with hydrogen. That the peculiar combustion of the taper is due to the great affinity of hydrogen for chlorine may be proved by a somewhat similar experiment, in which another body, composed only of hydrogen and carbon, and exceedingly rich in the former substance, is employed.

Experiment 124.—Saturate a piece of cotton wool or lint with *turpentine*, and then carefully remove the plate from a jar of chlorine and quickly introduce the wool into the jar.

Observe that the turpentine ignites spontaneously in the chlorine, producing a reddish flame, slight explosion, and dense clouds of smoke.

We learn from this that the body, turpentine, which is composed only of hydrogen and carbon, is decomposed by the action of chlorine—the chemical affinity of the

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chlorine for the hydrogen being so great that hydrochloric acid is formed and carbon liberated.

Coal gas for the same reason will burn in chlorine, with formation of hydrochloric acid and deposition of carbon.

Chlorine-water when exposed to sunlight will, in the same way, be decomposed by the action of the chlorine upon the hydrogen of the water, and in this case oxygen escapes.

Experiment 125.—Prepare some chlorine-water, as directed above, by means of the apparatus shown in Fig. 78; nearly fill a bottle with the solution, stopper it, and expose it to sunlight. After a few days examine the air in the upper part of the bottle by means of a glowing splint.

Observe that the fresh solution of chlorine in water has a distinctly yellowish tint, which it loses on standing, and that the gas which collects has the power to rekindle a glowing splint.

We learn from this that chlorine will decompose water with liberation of oxygen, in consequence of its great chemical affinity for hydrogen. For the reason that all organic substances contain hydrogen, for which body chlorine has a very great affinity, they are modified or destroyed by exposure to the gas.

Experiment 126.—Pass bubbles of chlorine through separate portions of water which have been coloured with (*a*) litmus solution, (*b*) red infusion of rose leaves, (*c*) sulphate of indigo.

Observe that in each case the colour disappears, leaving only a liquid with a pale yellow or brown tint.

We learn from this that chlorine acts as a bleaching agent, destroying organic colouring matters.

Experiment 127.—Introduce into a jar of *dry* chlorine a piece of *dry* Turkey red rag. After a time remove the red rag, moisten it with water, and reintroduce it into the jar.

Observe that the dry rag is not acted upon by the dry chlorine, but that when the body is moistened the red colour is discharged at once by the action of the chlorine.

We learn from this that chlorine is a bleaching agent which becomes active in the presence of water. The effect

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produced in this case is probably due to the action of the chlorine upon the water, in the first place, from which it liberates oxygen, which next converts the colouring matter into a colourless compound.

Chlorine will not only combine with non-metallic elements like phosphorus and hydrogen, but it has also a very great affinity for metals.

Experiment 128.—Remove the plate from a jar of chlorine and quickly introduce several pieces of *gold leaf*, "Dutch metal."

Observe that as the metal falls through the gas it burns spontaneously with a bright flame, and disappears.

We learn from this that chlorine has a great affinity for the metal, copper, for Dutch metal is an alloy which consists mainly of that metal. But this has only been taken as a type to illustrate the fact which has been demonstrated by many other experiments, that *chlorine will combine with all metals.*

Experiment 129.—Place a small piece of dry *sodium* in a deflagrating spoon, heat it in the Bunsen flame until it is quite melted, then immerse it in a jar of chlorine.

Observe that the sodium continues to burn, giving out a bright light and intense heat, and that white fumes are deposited.

We learn from this that chlorine has a very strong chemical affinity for sodium, and when they combine they form *common salt.*

The general deductions which we may draw from these typical experiments with chlorine and non-metals and metals is, that the body has a strong chemical affinity for both classes of elements.

So great is the chemical affinity of chlorine for hydrogen that it will take this body out of all compounds containing it except hydrochloric and hydrofluoric acids; since all

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organic bodies contain hydrogen, chlorine, as we have seen, possesses the power of decomposing them in consequence. It is also for this reason that chlorine is a valuable **disinfecting, deodorising** and **bleaching** agent.

Bleaching power of chlorine.—We saw above that we have good reason to believe that the bleaching property of chlorine in the presence of water depends upon the fact that it **liberates oxygen.** Ordinary free oxygen has very little, if any, bleaching power, the special activity of the oxygen therefore which is displayed in the act of bleaching by means of chlorine must be due to the increased power of the body at the instant of liberation. The oxygen in this state exhibiting abnormal properties is said to be **nascent.**

For manufacturing purposes large quantities of chlorine are employed in the form of **bleaching powder.** This substance is produced by the action of chlorine upon slacked lime. Quicklime is slacked with water and spread in layers on the floors of large chambers into which a current of chlorine is afterwards led; it is raked from time to time with long rods from the outside, and when the substance will no longer absorb chlorine it is known as **bleaching powder.** By the oilman this body is sold under the name of *chloride of lime.*

Bleaching powder is prepared for use by mixing with warm water and then decanting the clear liquid, into which the substance to be bleached is first dipped; the material is next rinsed in a bath of dilute sulphuric acid; this body liberates the chlorine from the solution of bleaching powder which has been taken up by the fabric. This operation is

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repeated until the required whiteness is obtained ; the cloth is then rinsed in a bath of clean water.

Experiment 130.—(*a*) Prepare a solution of bleaching powder with warm water and decant the clear liquid.

(*b*) Place in a second basin a very dilute solution of sulphuric acid.

(*c*) Into a third pour some clean water. Now introduce strips of coloured cotton, a few violets or other flowers, also a piece of printed paper, and one which has been written on with ordinary ink, into the liquids taken in the order *a*, *b*, *c*.

Observe that the colours fade in the case of the cotton, violets, and writing ink. Repeat this process until the colour is in each case entirely destroyed.

We learn from this that bleaching powder in the presence of an acid will destroy vegetable pigments, and that whilst it will bleach ordinary writing ink it does not discharge printers' ink.

Woollen and silk stuffs must not be bleached by chlorine, for the fibres are destroyed. We have seen that these materials may be bleached by the action of sulphur dioxide. Mineral colouring matters are not affected by chlorine ; nor is printers' ink, which contains much carbon, destroyed by this agent.

The chlorine which is required for the manufacture of bleaching powder is obtained by passing a mixture of hydrochloric acid gas and air over red-hot bricks impregnated with salts of copper. At a very high temperature the free oxygen is able to liberate chlorine, combining itself with the hydrogen and forming water. The hydrochloric acid gas, which is necessary in this operation, is obtained in enormous quantities as a bye-product in the manufacture of of soda from common salt; the first stage of which depends upon the action of sulphuric acid upon salt, by which means hydrochloric acid gas is liberated.

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SUMMARY TO CHAPTER X.

CHLORINE

- **Occurrence in Nature** — Chiefly as *chlorides of metals*, most abundant of which is *common salt;* sometimes present in volcanic gases, present in small quantities in plants and animals.
- **Properties** — Heavy greenish-yellow gas, does not burn in air, but will burn in an atmosphere of hydrogen; supports combustion of bodies rich in hydrogen, highly poisonous, soluble in cold water, powerful affinity for metals and non-metals; acts in presence of water as a bleaching agent; liberating oxygen indirectly it behaves as an *oxidising* agent.
- **Tests** — Distinguished by its colour and odour; it reddens, then bleaches, moist litmus paper.
- **Preparation** — By the action of sulphuric acid and manganese dioxide upon common salt. By the action of manganese dioxide upon hydrochloric acid.

BLEACHING POWDER

- **Composition** — Prepared by the action of chlorine upon slacked lime. It therefore contains calcium, oxygen, hydrogen, and chlorine.
- **Action** — Powerful bleaching action in presence of an acid.

QUESTIONS ON CHAPTER X.

1. How does chlorine occur in nature?

2. What are the chief properties of chlorine?

3. Describe fully how you would demonstrate the bleaching properties of chlorine.

4. How would you prepare chlorine, and how would you show that a jet of hydrogen burns in chlorine? Give a sketch of all the apparatus which you would employ.

5. Describe two processes for the preparation of chlorine. Sketch the apparatus in each case.

6. When paper moistened with turpentine is thrown into chlorine gas a flame is observed, accompanied by acid vapours and black smoke. Explain clearly what chemical change has occurred.

7. How would you prove that chlorine has a strong affinity for metals?

8. Why does chlorine water lose its colour when exposed to daylight?

9. You are required to prepare oxygen from chlorine water. Describe exactly how you will do it, and give a sketch of the apparatus which you propose to employ.

10. A jet of hydrogen is burnt in chlorine; what chemical change takes place?

11. Explain how chlorine acts as a bleaching agent, and how would you demonstrate this property of the gas?

12. How would you prove that common salt contains sodium and chlorine?

13. Describe the preparation of bleaching powder. How would you demonstrate the action of bleaching powder?

14. Explain the changes which take place when chlorine-water is exposed to sunlight.

15. What other bodies bleach besides chlorine?

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Preparation for Eleventh Lesson.

APPARATUS AND MATERIALS REQUIRED.

Hydrochloric acid, nitric acid, sulphuric acid ; sodic and potassic hydrates, litmus red and blue ; solid sodic carbonate, copper turnings, tinfoil, gold leaf, manganese dioxide, evaporating dishes, test tubes, tripod stand, sand bath, beakers, glass rods, and distilled water.

EXPERIMENTS TO BE PERFORMED.

" Samples of nitric and hydrochloric acids.

Show that both have all the properties belonging to acids, and that by neutralising them with potash and soda respectively nitre and common salt can be made.

Show the action of nitric acid on copper, tinfoil, etc.

Show that it has no action on platinum or on gold.

Copper placed in hydrochloric acid not attacked, but if mixed with manganese dioxide and warmed chlorine is given off."—*Science Department's Syllabus.*

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CHAPTER XI.

ACIDS.

All bodies which have sour taste, turn Blue Litmus red, and liberate Carbon Dioxide when added to solution of Sodium Carbonate—Sulphuric Acid has these properties—There are two other common bodies which have strong acid properties like Sulphuric Acid; these are Nitric Acid and Hydrochloric or Muriatic Acid; these are made of different constituents from Sulphuric Acid—All act on Litmus, etc., in the same way; all can be neutralised by Potash, forming Potassium Sulphate, or Nitrate, or Chloride—The compound formed by the union of an Acid and an Alkali is called a Salt—All three Acids are colourless liquids, but, beside the properties possessed by all acids, each acid has properties which belong to it alone—Nitric Acid attacks most metals—Poured on Copper the metal is dissolved and red fumes are formed—Hydrochloric Acid does not dissolve Copper, is not so heavy as Sulphuric Acid; when mixed with Manganese Dioxide and warmed a yellow irrespirable gas known as Chlorine is given off.

CHEMICAL COMPOUNDS may for convenience at this stage of our work be classified as *acids*, *bases*, and *salts*. We have seen types of the first of these in carbonic acid gas, sulphurous, sulphuric, and hydrochloric acids. They are all chemical compounds, for—

Carbonic acid contains carbon, oxygen, and hydrogen.
Sulphurous acid „ sulphur, „ „
Sulphuric „ „ „ „ „
Hydrochloric acid „ chlorine and hydrogen.

We saw in our lesson on oxygen that this body will combine with metals and non-metals—with the former to

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produce a class of substances which we shall now call bases, and with the latter to form acid-yielding bodies.

ACIDS are bodies which possess in a more or less marked degree the following properties :—

1. They are sour or acid to the sense of taste.
2. They change vegetable colouring matters from blue to red.
3. They cause effervescence when added to sodic carbonate through the liberation of carbonic dioxide.
4. They will neutralise alkalies forming salts.

All acids are further distinguished chemically by the fact that they contain hydrogen, which may be displaced by a base. These bodies may exist in any one of the three physical states. Thus tartaric acid is a solid, sulphuric acid is a liquid, and hydrochloric acid a gas.

There are two other very *common* and well-known bodies, in addition to sulphuric acid, which possess strong acid properties; these are nitric acid and hydrochloric acid.

Experiment 131.—Pour about 3 cc. of hydrochloric acid into an evaporating dish; then add about 10 cc. of water, and well stir the mixture with a glass rod. Touch the tip of the tongue with the moist glass rod. Dip a piece of blue litmus paper into the acid solution. Pour a little of the solution upon some solid sodic carbonate.

Observe that the solution has an acid flavour, that it reddens blue litmus, and that it causes the sodic carbonate to effervesce. The gas which escapes will turn lime-water milky.

We learn from this the two most important distinguishing properties of acids—viz., sour taste, and action upon blue litmus. These experiments should here be repeated with nitric and sulphuric acids.

It is usual to divide acids into two classes according to the presence or absence of oxygen. Those belonging to the first class are known as **oxyacids,** and those which do not contain oxygen are termed **hydracids.** Sulphuric, nitric, sulphurous, and carbonic acids are oxyacids. Hydrochloric and hydrosulphuric acids are hydracids; the former is composed entirely of hydrogen and chlorine, and the latter of hydrogen and sulphur.

Hydrochloric acid is prepared by the distillation of a mixture of **common salt** and **sulphuric acid;** a piece of apparatus, such as shown in Fig. 77, may be employed for

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this purpose. The body is a gas which is exceedingly soluble in water, and it is this solution of the gas which is known as *muriatic acid* or *spirits of salts*. When pure it is a clear liquid not nearly so heavy as sulphuric acid.

Nitric acid is prepared by heating a mixture of **nitre** and **sulphuric acid** in a retort. The acid which distils over may be condensed in a cold flask; when strong this body is highly corrosive; it destroys organic matters, and produces serious wounds if allowed to come in contact with the skin. It will dissolve most metals, and is in consequence known as *aqua fortis*.

Experiment 132.—Place some copper turnings in a test-tube, and in a second some tinfoil, and in a third a piece of *real* gold leaf or platinum foil. Add to each a small quantity of nitric acid.

Observe that the copper is at once dissolved by the nitric acid, with liberation of abundant red fumes and formation of a deep blue solution. The metallic tin disappears, but a white insoluble residue remains. The gold or platinum is quite unchanged.

We learn from this that nitric acid will dissolve copper. It will act upon tin, but does not dissolve the body, and it has no action upon gold.

Experiment 133.—Place in one test-tube some *copper turnings*, and in a second some *manganese dioxide*. Add to each a small quantity of strong hydrochloric acid.

Observe that the copper is not dissolved by the hydrochloric acid, but that on the addition of a little manganese dioxide and gentle warming, the copper dissolves and chlorine gas is evolved. If the mixture in the second tube is warmed chlorine gas is at once liberated.

We learn from this that hydrochloric acid will not dissolve copper, and that when mixed with manganese dioxide and warmed chlorine is evolved.

From these experiments we see that besides the properties possessed by all acids in common, each acid has properties which belong to it alone. Nitric acid would

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dissolve copper at once; hydrochloric acid, on the other hand, will not dissolve it. Neither nitric nor hydrochloric acid alone has any action upon gold, yet a mixture of the two acids will dissolve gold or platinum. This mixture is best prepared by adding 1 part by volume of strong nitric acid to 3 parts of concentrated hydrochloric acid; in consequence of its action upon the noble metals it is called **aqua regia.**

BASES.—All acids may be neutralised by bodies which are known as bases.

Experiment 134.—(*a*) Into an evaporating dish pour about 3 cc. of dilute hydrochloric acid, and into another basin introduce about 3 cc. of dilute nitric acid.

(*b*) Into two beakers pour about 3 cc. of strong potassic hydrate solution.

(*c*) To the acid solution in each evaporating dish add enough litmus solution to impart a decided colour.

(*d*) Then drop by drop, by the aid of a glass rod, pour some of the solution of potassic hydrate into each, until the blue colour is just restored.

Observe the gradual change in colour as the acids are neutralised.

We learn from this that both acids may be neutralised by a base. On evaporating the solution prepared by mixing potassic hydrate with nitric acid beautiful crystals of *nitre*, potassic nitrate, may be obtained (see experiment 30). If, in the second case, we substitute sodic hydrate for potassic hydrate, we may, on evaporating the solution prepared from the hydrochloric acid and sodic hydrate, obtain *common salt*. Bodies produced in this manner by the action of bases upon acids are known as **salts.**

The term **Base** includes all those substances which possess the power of partially, or entirely, neutralising the properties of a free acid. This property they have in consequence of the fact that they are all able to displace the hydrogen in the acid. They are divided into three groups according to their composition, viz.:—

(1) **Metallic oxides,** such as lime, oxide of iron, oxide of copper, etc.; some of these oxides are soluble in water, they then form substances which belong to the second class.

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(2) Metallic hydrates or alkalies, such as potassic hydrate, sodic hydrate, calcic hydrate, and baric hydrate.

(3) Certain compounds of hydrogen with nitrogen, phosphorus, carbon, arsenic, and antimony, the most important being a gaseous compound of nitrogen and hydrogen known as **ammonia.**

The **salts** which were produced in experiment 136 are named potassic nitrate, and potassic chloride or sodic nitrate and sodic chloride.

SUMMARY TO CHAPTER XI.

ACIDS

- **Properties** - Bodies which redden litmus, are sour to the taste, decompose sodic carbonate with liberation of carbon dioxide, neutralise bases. All contain hydrogen.
- **Classification** - *Oxyacids* contain oxygen; examples, nitric and sulphuric. *Hydracids* contain no oxygen; examples, hydrochloric and hydrosulphuric.
- **Nitric** - Prepared by distillation of nitre and sulphuric acid, will dissolve copper and act upon most metals, save gold, platinum, etc.
- **Hydrochloric** - Prepared by distillation of common salt and sulphuric acid, a gas which is very soluble in water; forms with nitric acid aqua regia.

SALTS

- **Preparation** - These bodies are produced by the action of a base upon an acid.
- **Properties** - They are all more or less saline to the sense of taste, usually neutral in reaction. They may be formed from hydracids or oxyacids, and so are oxysalts or haloid salts.

QUESTIONS ON CHAPTER XI.

1. How may chemical compounds be classified?
2. What is an acid? Give four examples of acids.
3. By what properties would you distinguish an acid?
4. How may acids be classified? Explain the terms hydracid and oxyacid, and give examples of each.
5. How is hydrochloric acid prepared?
6. What is aqua fortis, and how is it prepared?
7. How would you distinguish nitric acid from hydrochloric?
8. How would you distinguish sulphuric acid from hydrochloric?
9. Describe fully how you would neutralise any acid.
10. What is aqua regia, and how is it prepared?

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Preparation for Twelfth Lesson.

MATERIALS AND APPARATUS REQUIRED.

Specimens of caustic soda, caustic potash and ammonic hydrate, evaporating dishes, glass rods, blue and red litmus solution, apparatus for preparation of carbon dioxide, three large flasks provided with sound corks, large basin, beakers, specimens of salts.

EXPERIMENTS TO BE PERFORMED.

"Show that solutions of potash, soda, and ammonia turn reddened litmus blue, and that when a tube containing carbon dioxide is inverted in any of these solutions the gas is absorbed.

Add gradually to dilute sulphuric acid one of these bodies, and see that the acid character of the dilute sulphuric acid disappears.

Neutralise exactly sulphuric acid with potash, then evaporate and crystallise out the salt formed."—*Science Department's Syllabus.*

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CHAPTER XII.

ALKALIES.

A class of bodies which turn Red Litmus blue; have soapy taste and absorb Carbon Dioxide; if Potash be added gradually to Sulphuric Acid the properties of both bodies disappear, and at last a liquid is obtained that has no action on Litmus—This combination of acid and alkali is Sulphate of Potash or Potassium Sulphate; Sulphate of Soda or Sulphate of Ammonia can be formed in a similar manner.

ALKALIES belong to the class of bodies which in our last lesson we spoke of as **bases.** They include all those metallic oxides which dissolve in water, and the gaseous substance known as ammonia, which will also dissolve in water to form ammonic hydrate. They are distinguished as a class by the possession of the following properties :—

1. They turn red litmus blue; *i.e.*, restore the blue colour to reddened litmus.
2. They have a soapy taste.
3. They absorb carbonic acid gas.
4. They have a powerful chemical affinity for acids, with which bodies they unite to form salts.

Experiment 135.—(*a*) Place a piece of *caustic soda* about the same size as a small pea in an evaporating dish, and then add 10 cc. of distilled water and stir well with a glass rod until the whole of the solid has disappeared.

(*b*) Touch the tip of the tongue with the moistened rod.

(*c*) Rub a drop of the solution between the fingers, taking care to wash them thoroughly afterwards.

(*d*) Introduce a piece of reddened litmus paper into the solution.

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Observe that the solid readily dissolves in water, and the resulting liquid is soapy to the touch and taste, and that it turns litmus paper blue.

We learn from this the chief distinguishing properties of alkalies : namely, solubility in water, soapy touch, and power to restore blue colour to reddened litmus.

Experiment 136.—Pour into a beaker some dilute *sulphuric acid*, into a second introduce a small quantity of dilute *hydrochloric acid*, and in the third place a little dilute *nitric acid*. To each add sufficient litmus solution to impart a red colour. Next add to each drop by drop *sodic hydrate* solution.

Observe that on the addition of the alkaline sodic hydrate the red colour is gradually changed, and finally the distinct blue colour is restored.

We learn from this that the three common acids—viz., sulphuric, hydrochloric, and nitric—are neutralised by the addition of the alkali, sodic hydrate. These experiments may be repeated, employing *potassic hydrate* and *ammonic hydrate* in place of sodic hydrate.

Experiment 137.—Fill three long-necked flasks with *carbon dioxide*, which may be prepared as described in experiment 100. Into the first introduce a small quantity of a solution of caustic soda, the second caustic potash, and into the third pour some ammonia solution. Cork and shake each well. After a few seconds open each with its mouth immersed in water.

Observe that the water rises in the flasks, showing that the carbon dioxide has been absorbed.

We learn from this that **alkalies will absorb carbon dioxide,** and, on the other hand, we have seen in the last lesson that the common acids will liberate carbon dioxide from sodic carbonate.

Experiment 138.—Into an evaporating dish pour about 25 cc. of dilute sulphuric acid, colour the same with litmus solution. Prepare a strong solution of potash, and add this to the acid, drop by drop, with constant stirring, until the liquid is just neutralised, as in experiment 134. Evaporate one portion of the liquid nearly to dryness, and then allow it to cool. Pour out another portion into a dinner plate.

Observe that the sulphuric acid is neutralised by the potash, and that a stage is reached as the alkali is added to the acid, when at last a liquid is obtained that has no action on litmus. On heating and evaporating off

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some of the water a new body is obtained in the form of crystals; these have a decided **saline taste.**

We learn from this, once again, that by the combination of an alkali and an acid, here potash and sulphuric acid, a salt is produced ; in this case **sulphate of potash,** or **potassium sulphate.** In a similar manner, by neutralising portions of sulphuric acid with soda or ammonia, the salts, sulphate of soda and sulphate of ammonia, may be prepared.

SALTS are bodies which are formed when acids are partly or entirely neutralised by a base. As a rule these bodies have a pronounced **saline** taste, and they are more or less soluble in water. They form such an important class of substances, that it will not be out of place here to pause and try to learn something of their nature. We saw in last lesson that *every acid contains hydrogen.* Now this hydrogen of acid bodies may be in part, or wholly, displaced and replaced by a base ; the resulting body is in all cases a salt. From this we see that a salt may be described as a body in which the *hydrogen of an acid is partly or entirely replaced by a base.* From what we learnt of bases in our last lesson, we see that salts may be produced—

(1) By the action of **metals** upon acids.
(2) „ „ **metallic oxides** upon acids.
(3) „ „ „ **hydrates** „
(4) „ „ **ammonia** upon acids.

Since all the common acids will act upon metallic carbonates liberating carbon dioxide, the *salts of the common strong acids may be prepared by their action upon metallic carbonates.* But *carbonates are salts* themselves, and they may be produced by method 2 or 3 above ; for we have seen in experiment 137 that carbon dioxide combines directly with the alkalies soda, potash, and ammonia.

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1. In experiment 87, in the preparation of hydrogen, we saw that zinc acted directly upon sulphuric acid. If we concentrate the liquid which remains, we should obtain crystals of sulphate of zinc. This, then, is an example of the first method, mentioned above, of forming salts, for—

Zinc + Sulphuric acid = Zincic sulphate + Hydrogen.
A metal + *An acid* = *A salt* + *A gas.*

2. In experiment 105 we saw that carbon dioxide passed into lime-water formed the solid salt carbonate of lime; but carbon dioxide and water we saw formed carbonic acid, and lime is the oxide of the metal calcium which, with water, forms lime-water. Here then we have an example, from our past work, of the formation of a salt by the action of a metallic oxide upon an acid; for—

Lime + Carbonic acid = Carbonate of lime + Water.
Oxide of calcium + *An acid* = *A salt* + *A neutral body.*

3. We have many examples of the preparation of a salt by the action of metallic hydrates upon acids in our earlier lessons. Compare experiments 134 and 105, and see that in each case we have employed the oxide of a metal, which is soluble in water, to neutralise an acid. Here is an example :—

Sodic hydrate + Hydrochloric acid = Common salt + Water.
Sodium hydrate + *An acid* = *A salt* + *A neutral body.*

4. This method of forming salts was well illustrated in experiment 138, in which ammonia was made to combine with (*a*) hydrochloric acid, (*b*) carbonic acid, (*c*) sulphuric acid.

Hydrochloric acid + Ammonia = Ammonic chloride.
An acid + *An alkali* = *A salt.*

We have learnt that acids were divided into two classes —viz., those which contain oxygen, and those which do not contain oxygen. In a similar way the salts derived from these acids are arranged in two great groups.

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(1) **Oxysalts** are produced from *oxyacids*; for example, sodic, potassic, and ammonic *nitrates*, *sulphates*, and *carbonates* belong to this class.

(2) **Haloid salts,** yielded by the combination of bases with *hydracids*; we have examples of such in sodic, potassic, and ammonic *chlorides*.

Acids, in combining with bases to produce salts, may have part or the whole of the hydrogen which they contain displaced by the base, and in some few cases the quantity of bases which enters into the composition of the salt may be more than sufficient to take the place of the whole of the hydrogen present in the acid. It becomes necessary, therefore, to adopt another classification of salts, which depends upon the quantity of hydrogen which has been displaced from the acid by the base with which it has combined. The groups are as follows :—

1. **Normal salts,** or neutral salts, are those bodies in which the whole of the hydrogen of the acid has been replaced by a base.
2. **Acid salts** are those in which a portion only of the replaceable hydrogen of the acid has been displaced by a base.
3. **Basic salts** include all those which contain more of the base than is sufficient to account for the hydrogen which was present in the acids from which they have been derived.

NAMES OF ACIDS.—A few simple rules are adopted by chemists in naming acids, and these depend upon a knowledge of their composition ; thus :—

All **hydracids** have the prefix *hydro-* : for example, *hydrochloric acid* consists of hydrogen and chlorine ; *hydrobromic acid* is composed of hydrogen and bromine, and *hydrosulphuric* of hydrogen and sulphur. Observe that none of these contain oxygen.

Oxyacids are somewhat more difficult for the student to master the exact names of, for they vary with the proportion of oxygen present. Of two oxyacids containing the same chemical elements, the name of that which

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contains the *smaller proportion of oxygen* terminates in the syllable *-ous*, and that which contains the larger quantity of this element has the termination *-ic*, thus—

Sulphur*ous* acid contains less oxygen than
Sulphur*ic* acid.

If more than two acids contain the *same chemical elements*, but in different proportions, that which contains less oxygen than the acid which has the termination *-ous*, takes the prefix *hypo-*, and that which contains more oxygen than the *-ic* acid, has the prefix *per-*.

For example :—*hypo*-chlor*ous*, chlor*ous* chlor*ic* and *per*-chloric acids are all well known to the chemist.

Names of Salts.—The method adopted in naming salts is similar to that employed for acids, thus—

Salts of *hydro-* : acids take the termination *-ide*, sodic chloride.
„ *-ous* : „ „ „ *-ite*, „ sulphite.
„ *-ic* : „ „ „ *-ate*, „ sulphate.

It would be well now to study the names of the acids which we deal with in this "Chemistry of Common Objects," and in reading the following list the student should pause to think of the composition of each body.

Acids.	**Elements (contained).**			**Salts (formed).**
Sulphuric	*Sulphur*, *hydrogen*, and *oxygen*			**Sulphates.**
Sulphurous	„	„	„	**Sulphites.**
Nitric	*Nitrogen*	„	„	**Nitrates.**
Phosphoric	*Phosphorus*	„	„	**Phosphates.**
Silicic	*Silicon*	„	„	**Silicates.**
Carbonic	*Carbon*	„	„	**Carbonates.**
Acetic	„	„	„	**Acetates.**
Tartaric	„	„	„	**Tartrates.**
Stearic	„	„	„	**Stearates.**
Hydrosulphurous	*Sulphur*	„	—	**Sulphides.**
Hydrochloric	*Chlorine*	„	—	**Chlorides.**

The acids which contain carbon are so very numerous in animal and vegetable substances that quite distinctive names have been employed for them, as will be seen from the above list.

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SUMMARY TO CHAPTER XII.

BASES. Bodies which will neutralise acids, thus forming salts. **Metals, metallic oxides, alkalies,** certain *compounds* of **hydrogen** with **nitrogen,** etc., all behave as bases. **An alkali** is a base which is readily soluble in water; the more important are *ammonic*, *sodic*, and *potassic hydrates*. **Alkali metals** are those which form oxides which are easily soluble in water to form alkalies. **Alkaline reaction** is the turning blue of reddened litmus.

SALTS. Substances more or less **saline** to the taste, often **neutral** in reaction—that is, they will neither turn red litmus blue, nor blue litmus red. Those of **hydracids** are called **haloid salts.** Those formed from **oxyacids** are **oxysalts.** They may be normal, acid, or basic, according to the amount of base which has combined with the acid in the formation of the salt.

NEUTRAL BODIES. Substances such as water, which have *no decided saline, acid,* or *alkaline taste,* and which have *no action on red or blue litmus paper.*

QUESTIONS ON CHAPTER XII.

1. What is a base, and how are bases classified?

2. Describe an experiment by which you would demonstrate the formation of a salt.

3. How are salts classified? Give examples of each.

4. What is an alkali? Describe how you would demonstrate the chief properties of an alkali.

5. Explain the use of the terminations *-ic, -ous, -ide, -ite, -ate,* as employed by chemists.

6. Give examples of the use of the terms *alkali, alkaline, reaction, acid reaction,* and *alkali metal.*

7. What is a neutral body? Give an example to illustrate your answer.

8. Name the common acids and give the names of one salt of each.

9. How would you prepare a salt?

10. Explain the reason for dividing salts into two classes, *haloid* and *oxysalts.* Give two examples of each.

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Preparation for Thirteenth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Flasks, gas jars and plates, glass tubing, evaporating dishes, beakers, sand baths and tripods, glass rods, litmus solution and litmus papers, quicklime, hydrochloric acid, potassic hydrate, ammonic hydrate, sal-ammoniac, coal, mortar and pestle, horn, white of egg, cheese, glue, dried flesh, isinglass, smelling salts, test tubes, apparatus as shown in Figs. 81, 82, 83 and 84.

EXPERIMENTS TO BE PERFORMED.

"Prepare ammonia by treating ammonium chloride with an equal weight of slacked lime and enough water to make the whole into a thick mud: and demonstrate its smell, its action on red litmus, and its great solubility in water.

The gas passed into water, the increase of volume of the liquid. Its properties and their identity with those of the gas.

Volatility of ammonia shown by the liquid leaving no residue on evaporation.

Show that ammonium chloride is formed by neutralising a solution of ammonia with hydrochloric acid, and is obtained as a solid on evaporation, and that on further heating it is volatilised.

Heat coal in a coarse powder in a test-tube provided with a cork and delivery tube, and show that the liquid obtained is very alkaline.

Show the formation of ammonia by the addition of potash and lime to a solution of ammonium chloride."—*Science Department's Syllabus.*

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CHAPTER XIII.

AMMONIA.

A Gas with a very pungent smell—Solution in Water—One volume of Water dissolves 800 *volumes of Ammonia—This Liquid has the pungent smell of the Gas, and it can neutralise the strongest Acids—Formation of Ammonium Chloride or Sal-Ammoniac by Ammonia with Hydrochloric Acid—Ammonium Chloride a white solid, soluble in Water, with no smell of Ammonia—Ammonium Chloride a volatile body—The effect of boiling a solution of Ammonium Chloride with Lime or Potash —Ammonia is composed of* 82·3 *parts of Nitrogen and* 17·3 *parts of Hydrogen—The pungent odour of Smelling Salts is due to Ammonia—Animal matters, such as horn, dried flesh, glue, cheese, isinglass, heated so as to decompose these bodies, yield Ammonia—The formation of Ammonia in large quantities by heating Coal to make Coal Gas—Production of Ammonia when animal matters containing Nitrogen putrefy.*

AMMONIA is the name for the body which is commonly known as **Spirits of Hartshorn.** It belongs to the class of substances which we studied in our last lesson, for it is *one of the most powerful alkalies known.* A solution of the body was formerly prepared by the distillation of hartshorn, hence the name spirits of hartshorn. The name ammonia is derived from the fact that it was obtained from *sal-ammoniac,* a salt found in the earth near the temple of *Jupiter Ammon* in Libya. It is produced during the decomposition of animal and vegetable matters which contain nitrogen, and is yielded therefore by the distillation of glue,

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gelatin, white of egg, flour, cheese, horn, wool, and coal. It exists in minute quantities in the atmosphere, as we saw in Chapter V., and is dissolved by rain, by which means it is brought to the earth, where it serves as food for plants.

Ammonia is a colourless gas with a powerful, pungent odour, and strong acrid taste. It has been liquefied by cold and pressure, and is highly soluble both in water and alcohol. One volume of water will absorb about 700 volumes of the gas. Owing, therefore, to this extreme solubility in water, it must be collected by displacement of air or over mercury. The pure gas is composed of nitrogen and hydrogen, in the proportion of 14 parts by weight of the former to 3 parts by weight of the latter, or 100 parts by weight of ammonia contain 17·7 parts of hydrogen and 82·3 parts of nitrogen.

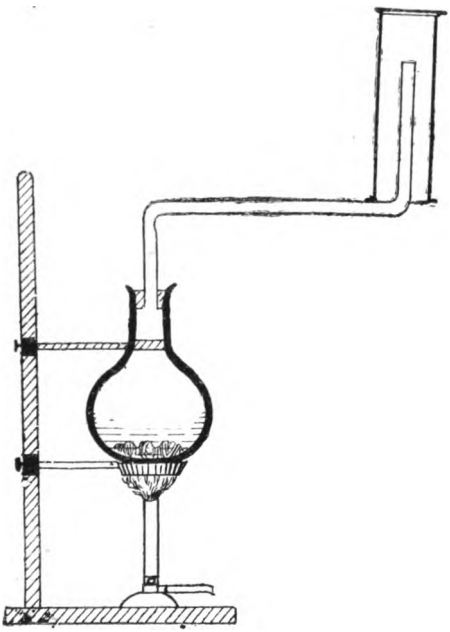

Fig. 81.—Preparation of ammonia gas.

Experiment 139.—Prepare a mixture of equal weights of *sal-ammoniac* and slacked lime, and make the whole into a thick paste with a little water. Introduce the substance into a flask, as shown in Fig. 81.

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Connect with the flask a delivery tube, bent twice at right angles. Adjust a jar for the collection of the gas by *upward displacement*. Heat the flask.

Observe that a colourless, pungent gas escapes which is lighter than air. The jar may be considered as full of gas when a piece of moistened red litmus paper brought to its mouth is turned blue.

We learn from this that ammonia may be prepared by heating a mixture of sal-ammoniac and quicklime, that it is invisible, has a strong odour, and is lighter than air.

The substances sold under the names "**smelling salts**" and "**sal-volatile,**" which *smell so strongly of ammonia*, contain carbonate of ammonium. The odour is due to the fact that they are volatile and give off ammonia gas when exposed to air.

Experiment 140.—Plunge a burning taper into a jar of ammonia.

Observe that the flame is at first somewhat enlarged near the mouth of the jar, and then is extinguished by the gas.

We learn from this that ammonia will not support combustion, but that when its temperature is somewhat raised by the taper it will burn in air. If the end of the delivery tube from which the ammonia is escaping is heated, the hot gas may be ignited, and it then burns with a pale, greenish yellow flame.

Experiment 141.—Introduce the closed mouth of a jar of the gas into water which has been stained with red litmus solution.

Observe that the water rises rapidly to fill the jar, and that the litmus turns blue.

We learn from this that ammonia is very soluble in water, and that it has an alkaline reaction. For lecture purposes this fact about ammonia may be strikingly illustrated by means of the apparatus shown in Fig. 82. Fit up two flasks, as in the illustration. The cork, *C*, of the lower flask, *A*, is provided with two holes; through one of these a long glass tube passes up into the flask, *B*, the

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cork of which has only one hole. The second hole of *C* is fitted with a short delivery tube, *D*. The upper flask is filled with ammonia, and the lower with water coloured red with litmus solution. The flasks should be placed in position, as shown, and then a long piece of indiarubber tubing should be connected with *D*, and a little water should be urged up into *B* by blowing through *D*. The water blown into *B* absorbs the ammonia, and the liquid flows from *A* into *B*, changing in colour from red to blue.

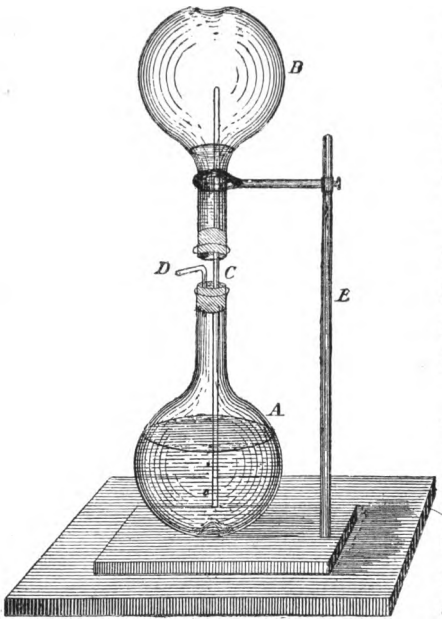

Fig. 82.—Ammonia soluble in water.

One part of water by volume will absorb 1050 volumes of the gas at 0° C., and at the ordinary atmospheric temperatures one volume of water will absorb from 700 to 800 volumes of ammonia. When the gas is dissolved by water there is a distinct increase in the bulk of the liquid; hence, bulk for bulk, a strong solution of ammonia must be lighter than water. The strong solution has a specific gravity of ·880, and is known as *liquor ammoniac fortior*. From this body ammonia gas is freely liberated on heating.

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A piece of apparatus is fitted up, which consists of a piece of wide glass tubing, bent to form the letter W, and provided with a well-fitting cork at *a*, and a cork and delivery tube *r*, at *e*. Into the bend *b*, a small quantity of *strong ammonia solution* is poured, and the limb *d* is loosely packed with *granular quicklime*. The tube is heated at *b*, ammonia is then liberated from the hydrate and dried by the quicklime in *d*, and escapes by *r*.

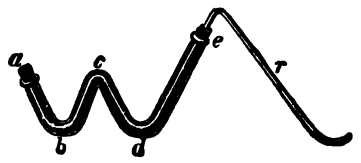

Fig. 83.—Apparatus for preparation of ammonia gas from ammonia hydrate.

Experiment 142.—Pour some strong ammonia solution into an evaporating dish placed on a sand bath, supported on a tripod stand, and heat the liquid.

Observe that the whole of the substance is evaporated, nothing remains.

We learn from this that ammonia hydrate is a *volatile* body.

Chloride of ammonia.—In experiment 24, we saw that ammonia gas and hydrochloric acid gas combined directly together to form a new substance, which was a white opaque snow-like solid; this compound is commonly known as *sal-ammoniac* or *ammonium chloride*.

Experiment 143.—Place in an evaporating dish some *ammonia solution*, and add, drop by drop, *hydrochloric acid* until the liquid is neutral. Place the dish on a sand bath, and *carefully* evaporate.

Observe that a white solid body is produced by the combination of hydrochloric acid and ammonia, without odour, and with a saline taste. Introduce a little of this into water, and notice that it dissolves at once.

We learn from this that the salt, ammonium chloride, is a whitish, inodorous, saline, soluble, solid body, produced by the neutralisation of ammonia with hydrochloric acid.

Experiment 144.—Introduce some of the solid sal-ammoniac prepared in experiment 143 into a crucible, and apply heat.

Observe that white fumes are given off, and that the solid disappears.

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We learn from this that the solid ammonium chloride is a volatile body.

Experiment 145.—Pour a solution of ammonium chloride into a test-tube, and add a few pieces of quicklime or solution of potash, and boil.

Observe that ammonia is given off; which body may be identified by its odour and action upon red litmus paper.

We learn from this that ammonia gas is liberated from ammonium chloride by the action of lime or potash.

That the *destructive distillation of nitrogenous organic compounds* is attended by the liberation of ammonia may be illustrated by the following experiments.

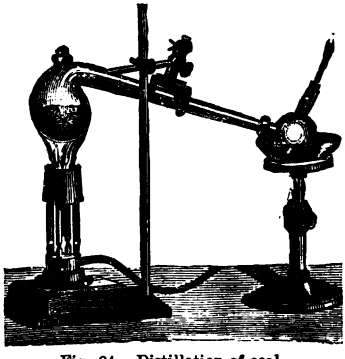

Fig. 84.—Distillation of coal.

Experiment 146.—Into a test tube introduce a few pieces of *gelatin, cheese, glue, white of egg, wool, dried flesh* or *horn*. Place the tube in a horizontal position on a sand bath, and apply heat. Test the gas which escapes by means of red litmus paper.

Observe that the mass when heated chars, that steam is given off, and that the red litmus is turned blue. In most cases there is also a decided ammoniacal odour.

We learn from this that ammonia is evolved when certain organic bodies undergo decomposition. The ammonia can be obtained from any of these bodies with greater ease, and at a lower temperature, without smoke and fumes, if the substance used be first reduced to a powder, and mixed with soda-lime or quicklime.

When such organic substances as those mentioned above naturally rot, their decomposition is attended by the liberation of ammonia, and it is in part from this source that the atmospheric ammonia is derived. During the combustion of coal, and in the manufacture of coal gas, ammonia is produced.

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Experiment 147.—Fit up a hard glass retort with a bulb at the end of its tube. Place in the retort some powdered coal and apply heat.

Observe that a tarry matter collects in the bulb, and a gas escapes which will burn with a luminous flame, and that carbon remains in the retort. Test the liquid which collects in the bulb with red litmus; notice that it is turned quite blue. See Fig. 84.

We learn from this that during the destructive distillation of coal the volatile alkali, ammonia, is given off.

SUMMARY TO CHAPTER XIII.

AMMONIA

- **Occurrence in Nature** — Traces exist in the air, otherwise it occurs in combination with acids as salts. The waters of some rivers contain *ammoniacal salts* in small quantities.
- **Properties** — A colourless gas with a strong pungent odour; non-supporter of combustion, but when heated it burns with a greenish flame. Extremely soluble in water, with which it enters into combination, forming *ammonic hydrate*, which behaves in many respects like the hydrates of sodium and potassium. It combines with acids directly to form salts.
- **Tests** — Ammonia may be detected by its strong pungent odour by its action on red litmus, and by the heavy white fumes which it yields with hydrochloric acid gas.
- **Composition** — Ammonia consists of hydrogen and nitrogen combined together in the proportion of 3 parts of the former to 14 parts of the latter by weight.
- **Preparation** — By heating *nitrogenous compounds* of organic origin with quicklime or soda lime.
  By heating a mixture of *sal-ammoniac* and quicklime.

QUESTIONS ON CHAPTER XIII.

1. What are the chief sources of ammonia?

2. How would you prove that gelatine yields nitrogen?

3. How would you prepare ammonia gas and fill a jar with it? Give a sketch of the apparatus.

4. Ammonia gas is led into a solution of hydrochloric acid. What is produced? The resulting liquid is evaporated to dryness in a porcelain crucible, which is afterwards heated to redness. What will remain in the crucible?

5. In what respect does water saturated with ammonia differ from pure water?

6. Sal-ammoniac and lime are mixed together and heated in a flask. What are the properties of the gas given off, and how can they be shown? Give sketches.

7. What takes place when a lighted taper is immersed in gaseous ammonia?

8. Show how you would prove that ammonia will combine with acids, forming salts.

9. Say what you know of sal-ammoniac and spirits of hartshorn.

10. How would you prove that ammonia is produced during the distillation of coal?

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Preparation for the Fourteenth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of limestone, marble, chalk, bones, coprolites, corundum, alum, quartz, sand, flint, quicklime, Iceland spar, hydrochloric acid, lime-water, litmus red and blue, apparatus for generation of carbon dioxide as in Fig. 87, kaolin, pipe-clay, fire-clay, Fuller's earth, ordinary brown clay, glazed and unglazed earthenware, red and white ware, crucible, tripod, flasks, beakers, and test-tubes.

EXPERIMENTS TO BE PERFORMED.

"Show that limestone, chalk, marble, and oyster-shells effervesce with dilute hydrochloric acid, and that a gas, carbon dioxide, is evolved.

Heat a piece of limestone or marble to redness in a fire, and show that after heating it no longer gives off carbon dioxide on treatment with an acid.

Show that a piece of moistened red litmus paper pressed against limestone is not affected, but that when pressed against lime it is turned blue.

Show slacking of lime; draw attention to heat evolved. No such result on treating limestone with water.

Show that lime is soluble in water, whereas limestone is not.

Add carbon dioxide to the solution of lime, and show that white powder is formed which on treatment with acid evolves carbon dioxide again.

Show plasticity of clay, and exhibit one or two specimens of ware before being baked.

Show that a vessel of kneaded or "puddled" clay will hold water.

Show alum, and demonstrate that alumina is contained in it by heating ammonia alum.

Show specimen of aluminium, and explain that this metal is contained in alumina and therefore in clay."—*Science Department's Syllabus.*

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CHAPTER XIV

LIME AND CLAY.

Limestone, Marble, Oyster-Shells and Chalk, all contain a Metal known as Calcium—The Oxide of this Metal known as Lime—Lime and Carbon Dioxide are together present in Limestone, Marble, Shells, and Chalk—When these are strongly heated, especially in a current of Air, the Carbon Dioxide is evolved and the Lime is left—Action of Water on Lime—Use of Lime in making Mortar—Lime slightly soluble in Water—On blowing Carbon Dioxide into a clear Solution of Lime (Lime-water) Liquid becomes turbid, owing to combination of Lime with Carbon Dioxide to form Chalk—Same effect on breathing through Lime-water—Other important Salts of Lime are Gypsum or Plaster of Paris (Sulphate of Lime) and Phosphate of Lime, which latter exists largely in Bone—Clay is a combination of a body called Silica, which is the chief constituent of Sand and Flint, with the Oxide of a Metal known as Aluminium, so called because it exists in Alum—Glass is a compound of Silica with Lime and an Alkali, Potash or Soda—Varieties of Clay; their use in manufacture of Bricks and Pots—The Metal contained in Clay (Aluminium), a white body with a brilliant lustre, 2½ times as heavy as Water; may be rolled out into thin Sheets and drawn into fine Wire, not oxidized in the Air.

LIME is a very abundant substance in the earth, for it enters into the composition of all limestone, marble, chalk, coral, shells, and the skeletons of all animals. In nature it is never found free, but always combined with other substances. The body with which it is united in all of the above materials, and with which it is most frequently found combined, is carbon dioxide. **Carbonate of lime** is by

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far the most abundant ingredient in all the above-mentioned bodies. Lime is a chemical compound of **oxygen**, and a metal known as **calcium**. We see then that carbonate of lime must therefore contain the chemical elements carbon, oxygen, and calcium ; for it is made up of carbon dioxide and calcium oxide or lime. Carbonate of lime is met with in the earth in small quantities in the pure state as a beautiful crystalline mineral known as **calcite** or **Iceland spar.**

Lime may be obtained from the natural carbonate by the application of heat, and it is in this way that men prepare lime from chalk and limestone by roasting the limestone in **kilns.** By this means the carbon dioxide is expelled and lime remains. From this lime it is possible, but very difficult, to extract the metal calcium, and oxygen ; and it has been shown by experiments that 56 parts of lime by weight contain 40 of calcium and 16 of oxygen. **Calcium** is a soft yellowish metal, which has a very powerful affinity for oxygen, and it is in consequence of this latter property that the metal is exceedingly difficult to extract from lime.

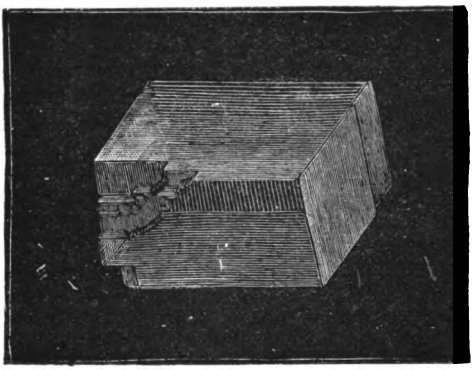

Fig. 85.—Crystal of calcium carbonate.

Lime when pure is a white infusible solid, which gives out a bright light known as the **limelight** when intensely

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heated. It combines very readily with water, increasing in bulk, and giving out much heat during the combination. After a time the body produced falls to a powder, which is known as **slacked lime** (see experiment 23). In slacking, 56 parts by weight of lime will unite with 18 parts by weight of water, and the substance produced is a perfectly dry powder. This *slacked lime*, *hydrated lime* or *calcium hydrate* as it is variously called, is slightly soluble in water. The clear solution is the liquid which we have so often employed in our past experiments under the name of **lime-water.** Freshly slacked lime mixed with sharp clean sand into a paste sets into a hard mass on standing. This body is largely employed in building under the name of **mortar.**

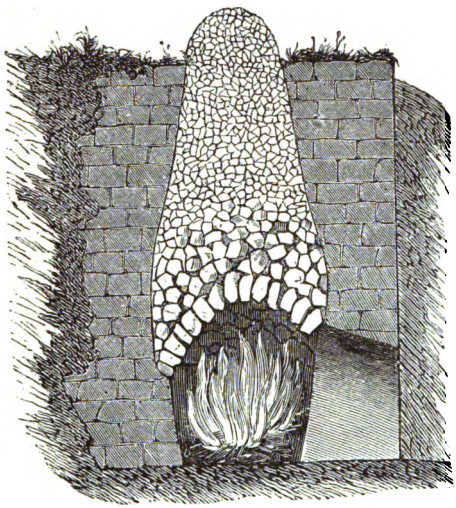

Fig. 86.--Preparation of quicklime.

On passing carbon dioxide into clear lime-water (see experiment 105), the liquid becomes turbid owing to the combination of lime with carbon dioxide to form carbonate of lime. The same effect is produced, as we have seen in experiment 54, by breathing into lime-water. Very carefully performed experiments have shown that 44 parts of

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carbon dioxide always combine with 56 parts of lime, and form 100 parts by weight of carbonate of lime.

Since 44 *parts of carbon dioxide contain* 12 *of carbon and* 32 *of oxygen,*
and 56 „ *lime (calcium oxide)* „ 40 *of calcium* „ 16 „
„ 100 „ *carbonate of lime* „ 40 „ „ 12 *of carbon*
and 48 *of oxygen.*

Other salts of lime occur largely in the earth. Of these the most important are **gypsum,** from which plaster of Paris is made, and **coprolites** and **apatite,** which are of value for manures. Gypsum and alabaster consist of **sulphate of lime,** whilst coprolites, apatite, bones, and shells contain much **phosphate of lime.**

Experiment 148.—Place in test-tubes pieces of *limestone, marble, chalk, coral,* and some broken *oyster-shells.* Add to each of the solids a few drops of dilute hydrochloric acid.

Observe that effervescence occurs, and that the gas which escapes renders lime-water milky, and extinguishes a burning taper.

We learn from this that these bodies all give off carbon dioxide under the influence of hydrochloric acid.

Experiment 149.—Heat a piece of chalk or limestone in the bright red part of a fire. After it has become quite red hot, remove it from the fire, allow it to cool, and then place it on a plate and pour on to it some dilute hydrochloric acid. The limestone or chalk may be heated instead in an open crucible by means of the blowpipe flame.

Observe that no effervescence now occurs when the acid is added to the roasted limestone, and that, therefore, no gas is evolved.

We learn from this that by heating limestone to redness the carbon dioxide is driven off and **lime remains.** This plan is followed in the preparation of **quicklime.**

Experiment 150.—Press pieces of moistened red and blue litmus papers against fragments of chalk or limestone, and quicklime.

Observe that the quicklime turns the red paper blue, but that in the other cases the papers are not acted upon.

We learn from this that *quicklime is alkaline in reaction,* but that chalk and limestone are neutral.

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Experiment 151.—Place on one dinner plate a block of quicklime, and on another a block of marble. Slowly pour on to each some water.

Observe that the quicklime remains dry, clouds of steam are evolved, and the mass swells, cracks and falls to a powder. No such changes are produced by the addition of water to any one of the other bodies mentioned.

We learn from this that quicklime combines with water, and a great amount of heat is developed by the chemical combination of the two bodies; but limestone, marble, etc., are not so acted upon by water.

This **slacking of lime** by the addition of water is a preliminary part of the operation of mortar making.

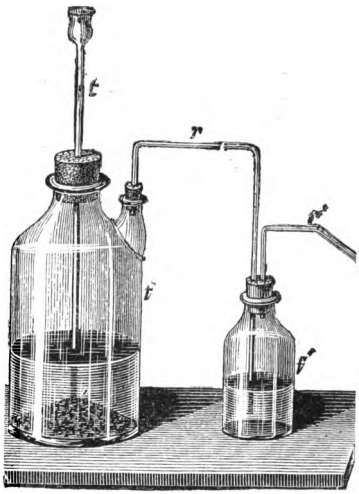

Fig. 87.—Apparatus for generating carbonic acid from limestone.

Experiment 152.—Place some of the powdered slacked lime prepared in the last experiment in a large flask, add sufficient distilled water to two-thirds fill the flask, and shake well; allow to stand, and after a time decant the clear liquid.

Observe that the water now has an alkaline reaction and taste.

We learn from this that lime is soluble to a limited extent in water, and that the resulting liquid is an alkali. If limestone be treated in the same way it does not dissolve.

Experiment 153.—Fit up an apparatus, as shown in Fig 87, for the generation of carbon dioxide. Place in the Woulffe's bottle on the left some fragments of broken marble. Into the second bottle pour some clear lime-water. Adjust all corks and tubes, and then pour some dilute hydrochloric acid into the thistle funnel, *t*.

Observe that the marble effervesces on the addition of the acid, and a gas is evolved which bubbles into the lime-water and turns it milky. After a time decant the liquid from the second bottle, and collect and dry the

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white precipitate; add to a portion a few drops of acid, and notice that it effervesces and carbon dioxide is evolved.

We learn from this that the powder deposited from the turbid lime-water, which has been produced by the combination of carbon dioxide and lime, is identical in composition with pure limestone, for it consists of lime and carbon dioxide.

Chalk and limestones and the other naturally occurring compounds which contain lime, as carbonate of lime, also contain other bodies. For example, the purest chalk contains about nine parts in one hundred of silica. So that although in experiment 153 we have *prepared carbonate of lime* by strictly chemical means, *it is not chalk*, for this body contains small and variable quantities of other ingredients. At the same time we must bear well in mind the fact that carbonate of lime is by far the most abundant ingredient in marble, limestone, and chalk.

Experiment 154.—Recharge the apparatus which was employed in the last experiment, and allow the carbon dioxide to pass for a much longer time into the second bottle, which contains the lime-water.

Observe that the lime-water which is first turbid becomes in time clear.

We learn from this that the *carbonate of lime* which is **insoluble** is converted by an *excess* of *carbonic acid* gas into some new body which **redissolves** into the water; this new substance *containing twice as much carbon dioxide combined with the lime* is known as **bicarbonate of lime** or **calcium bicarbonate.** If a little of the clear liquid obtained at the end of this experiment is boiled in a test-tube, a white deposit is produced of carbonate of lime; this body when tested with hydrochloric acid dissolves with violent effervescence and liberation of carbon dioxide.

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By heating the liquid containing the *soluble calcium bicarbonate* the excess of carbon dioxide is expelled and the *insoluble calcium carbonate* is reformed and precipitated. Although the carbonate of lime, of limestone, marble, chalk, etc., is insoluble in water, yet rain and ground water holding carbonic acid gas in solution are able to dissolve out some of the lime as bicarbonate of lime. The rain gains carbonic dioxide during its passage through the air, and ground water gains it from the decomposing organic matters of soils, and such waters by reason of the excess of carbonic acid gas which they bring into contact with carbonate of lime possess the power of converting it into soluble bicarbonate of lime. Waters holding bicarbonate of lime in solution are hard, and on heating yield a **fur** in the kettle or an **incrustation** in the boiler.

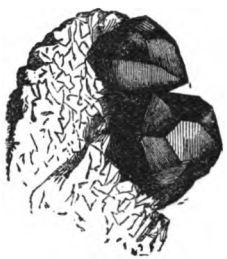

Fig. 88.—Rubies (*oxide of aluminium*).

CLAY is exceedingly abundant in the earth, and it is met with in various forms. It is always composed of exceedingly fine particles which have been worn down from various rocks by water and then deposited. Clay is rarely found pure, but is usually mixed with various substances which impart to it characteristic properties; generally it occurs as a soft plastic body which can easily be moulded into various shapes; but it also exists in a highly changed and hardened form as **slate.** In a moist condition it is **impervious** to water, but able to hold a considerable quantity by mechanical adhesion; on drying it shrinks considerably, and if strongly ignited parts with all its water and forms a porous, infusible, brick-like mass.

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Clay is chemically composed of two oxides, just as carbonate of lime is made up of two oxides. The two bodies which are united to form clay are **silica** and **alumina**; so that pure clay is **silicate of alumina.**

Silica exists free as **quartz, flint,** and **sand**; it is the oxide of the non-metallic body **silicon.** Silica is largely employed in the form of sand in the preparation of **glass.**

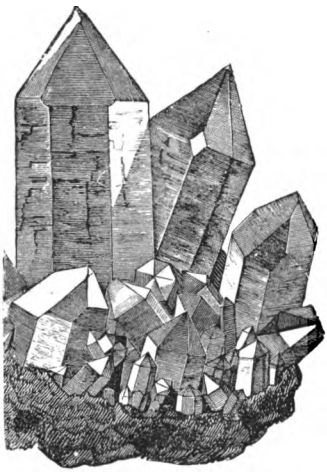

Fig. 89.—Quartz crystals (*oxide of silicon*).

Alumina is the oxide of the metal **aluminium**; this metal is not met with free in nature, but the oxide, alumina, occurs in the earth as **corundum, ruby, sapphire,** and **emerald.** Aluminium derives its name from the fact that it can be obtained from **alum**, which is a **sulphate of aluminium** and **potassium.**

The two bodies, oxide of silicon and oxide of aluminium, are combined together to form silicate of alumina, which is another name, as we have seen, for pure clay.

The numerous varieties of clay contain impurities which impart to them special properties. Among the purest forms of clay we have—

(*a*) **Kaolin** or **China clay,** a fine white clay which is used in the manufacture of china.

(*b*) **Pipe clay,** a white plastic body which is employed in the manufacture of tobacco pipes and white pottery.

(*c*) **Fuller's earth,** yellow, red, blue, and purple clays containing varying percentages of iron and other impurities.

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(*d*) **Fire clay** is rich in silica, and yields infusible bricks.

The **impure varieties of clay** are all of more or less value for agricultural purposes; some of the heaviest clays when well drained and cultivated furnish the richest soil. They are also of value for making brown earthenware, drain pipes, tiles, flower pots, bricks, etc.

Experiment 155.—Knead some kaolin or ordinary clay with water, mould it into the form of a basin, and show that it will hold water.

Observe that the clay is soft and greasy to the touch, plastic, and may easily be moulded, and that it will hold water. Place the vessel on one side to dry, and notice that it becomes quite hard.

We learn from this that clay is a soft plastic body which will hold water. For this reason canals and ponds are lined with clay. For the manufacture of pottery the clay is beaten up in water till the fine parts are suspended in the liquid, which is then passed through large hair sieves. The fine solid powder is then collected and dried in a kiln, and after being reduced to a proper consistency by mixing with water, it becomes fit for the manufacture of cups, dishes, and other kinds of hardware.

The smoothness and gloss of earthenware and china are produced by first baking the clay goods for a time to harden them; they are then dipped into a vessel containing a thin mud of finely-powdered quartz and felspar, and are next strongly heated, so as to melt this coat, which fuses and forms a colourless glass. This glazed surface enables the vessels to hold liquids without allowing any to ooze through, and it resists the action of acids and alkalies.

Alumina, the body which is combined with silica in clay, may be prepared by heating **ammonia alum** in a crucible. Ammonia alum is a double sulphate, consisting of sulphuric acid in combination with the two bases, ammonia and aluminium; the former of which is driven off by the application of heat.

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Experiment 156.—Place some powdered *ammonium alum* in a crucible, fix it on a wire triangle supported on a tripod stand, and heat it strongly by means of a Bunsen flame.

Observe that the substance froths up and ammonia escapes, and a fine white solid body remains. Allow this to cool, and observe that it has a dry earthy taste.

We learn from this that ammonia alum may be decomposed by heat, and thus alumina obtained. Alumina, as explained above, is a compound of aluminium and oxygen.

ALUMINIUM is a metal with a brilliant white colour. It takes a high polish, and does not tarnish; nor does it oxidize in air even when heated, but it may be burnt in oxygen. Aluminium is very tough and light, and can be forged or worked into almost any shape. It is remarkably resonant when struck. In consequence of these many valuable properties its use is rapidly extending in the arts. It is interesting to remember that although this element does not occur free in nature, yet it is the **most abundant of all metals.** Combined with other elements it is found entering into the composition of an enormous number of minerals. Of the three elements of which the crust of the earth is principally composed—viz., oxygen, silicon, aluminium—it ranks third in order of quantity.

GLASS.—Silica exists largely in the free state in nature, and as sand it is employed in enormous quantities in the manufacture of glass. Silica heated alone is almost infusible, as is lime also; but when mixed the two readily fuse together, forming **calcium silicate.** Glass may be regarded as an **artificial silicate,** which contains silica as its acid oxide, but in which the base is variable. *There are four principal kinds of glass.*

1. Flint Glass, so called because flint is employed in its preparation, is a silicate of potassium and lead; it is used for optical instruments and cut glass ware. It is highly lustrous, and when modified in composition by the addition of borax it is employed for paste or imitation diamonds.

2. Window or **Crown Glass** is a silicate of sodium and calcium.

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3. Bottle Glass is a mixture of silicates of sodium, calcium, aluminium, iron, etc.

4. Bohemian Glass is silicate of potassium and calcium; it is not easily fused, and hence is employed for chemical apparatus where high temperatures are required.

SUMMARY TO CHAPTER XIV.

LIME

- **Occurrence in Nature** — Never free or uncombined, but chiefly as *carbonate of lime*; as such it enters largely into the composition of such bodies as marble, limestone, chalk, bones, and shells. It also exists as *sulphate of lime* in gypsum and *phosphate of lime* in bones.
- **Preparation** — Lime is obtained from the carbonate by heating in kilns, by which means the carbon dioxide is driven off.
- **Properties** — Lime is a white, infusible solid, which combines with water to form *calcium hydrate* or *slacked lime*. This is somewhat soluble in water, with which it forms lime-water, which has an alkaline reaction.
- **Composition** — Lime is composed of the metallic element calcium and the non-metallic element oxygen. These form calcic oxide, which combined with carbon dioxide forms carbonate of lime.

CLAY

- **Occurrence in Nature** — Exists in many forms as rocks. The purest is white, and the various coloured varieties are due to impurities.
- **Composition** — Composed of *silica* in combination with *alumina*. Silica exists free as *sand*, *flint*, *quartz*, etc. Alumina is found in small quantities free, as *corundum*, *sapphire*, *ruby*, and *emerald*. *Silica*, oxide of silicon. *Alumina*, oxide of aluminium. These two oxides combined form *silicate of alumina*, the basis of clay.

QUESTIONS ON CHAPTER XIV.

1. What is lime, and how does it occur in nature?

2. How is lime obtained from limestone?

3. What other substances contain lime besides limestone and chalk?

4. What is the chemical name for the compound of lime present in chalk? What elements does it contain? What other substances contain this body?

5. How would you demonstrate the chief properties of lime?

6. How would you distinguish limestone and chalk from quicklime?

7. Describe how you would prepare lime-water.

8. Give an account of the chemical composition of clay.

9. How would you demonstrate the chief properties of clay?

10. How do silica and alumina occur in nature?

11. Name some different kinds of clay, and explain the uses to which they are put. What are the causes of the difference?

12. How may alumina be prepared from alum?

13. What is glass? Give an account of the different kinds of glass.

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Preparation for Fifteenth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of all the *common metals*, and sulphides, chlorides, and oxides of the same; specimens of the following alloys: brass, pewter, bell-metal, gun-metal, speculum metal, German silver, type-metal; crucible, tripod stand, wire triangle, bismuth, tin, lead, cadmium, balance, test-tube, and thermometer.

EXPERIMENTS TO BE PERFORMED.

Show solubility of certain metallic oxides in water; for example, soda, potash, and lime.

Demonstrate that metals dissolve in acids to form salts.

Prepare *fusible metal*, and demonstrate its chief properties.

Prepare some sodium amalgam, and demonstrate its chief properties.

Show how to remove mercury from the surface of gold and silver.

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CHAPTER XV.

METALS.—INTRODUCTORY.

About seventy different Elementary Substances known—All the common Metals are Elements; for instance, Iron, Lead, Copper, Zinc, Mercury, Silver, Gold, and Tin are Elements—All combine with Oxygen to form Oxides, with Chlorine to form Chlorides, and with Sulphur to form Sulphides.

In Chapter III. we saw how chemists classify all kinds of matter under the two heads, **elements** and **non-elements.** The **chemical elements** are the simplest forms of matter known. Further, we learnt that about **70 different elementary substances** have been discovered. Of these *some are very abundant*, like *oxygen*, *silicon*, and *aluminium*, *others are rare*, as *gold* and *platinum*, which exist only in very small proportion in the earth. Some, like *calcium*, *aluminium*, *chlorine*, and *hydrogen*, are *never met with* free, whilst others are found native in nature; for example, *nitrogen* and *oxygen* in air, and *sulphur* and *carbon* in the earth. These various chemical elements, whether they occur free or are only found in combination with other bodies, are arranged in two classes—as **non-metallic elements** and **metallic elements**, and of the seventy known elements the majority are metals. Our past lessons have been devoted to the

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study of the properties of the more important and abundant non-metallic elements, and the compounds which they unite to form. Thus we have learned something of the preparation and properties of **oxygen, hydrogen, carbon** and **nitrogen,** which are by far the **most abundant constituents of all living things.** We have studied sulphur and chlorine, the two elements which, after oxygen, are found most plentifully in the earth in combination with the metallic elements. The properties and composition of many of the more important inorganic compounds containing these elements have also been studied. The student should now know something of the chemistry of air, water, carbonic acid gas, ammonia, coal, lime, clay, sand, earthenware, glass, and the common acids and alkalies.

The next five chapters we shall devote to the study of the important properties of the **common metals,** including in our work, in the first place, all those which are most abundant in the rocks of the earth and in the ashes of living things, and therefore those, a knowledge of which is important to the student of Agriculture, Geology, Physiology, and Physiography. In the second place we shall treat of those which are of the greatest industrial importance. We must now consider the chief facts about the great class of metallic elements.

METALS are distinguished by certain well-marked physical and chemical properties which they possess in varying degrees, and by means of which they may be identified with more or less ease.

1. At ordinary temperatures they are *solid bodies of high specific gravity,* whilst the non-metallic elements are, under the same conditions, either

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gases or solids of low specific gravity. These are distinctive properties, but no very hard-and-fast line can be drawn; for of the metals *mercury* is a liquid, *lithium* is the lightest known solid, sodium and potassium will also float on water. Of the non-metals *bromine* is a liquid, and *iodine* yields the heaviest vapour known.

2. The clean surfaces of metallic elements reflect light in a peculiar way, in consequence of which they are said to possess *metallic lustre*. The solid non-metals, on the other hand, have different lustres; they are *glassy*, *resinous*, *dull*, *silky*, etc.

3. They are usually *tenacious*, *malleable*, and *ductile*; they may be beaten into thin plates or drawn out into fine wires. The solid non metallic elements are more or less brittle. We have seen a good example of such in the case of the element *sulphur*.

Fig. 90.—Crystals of blende (*sulphide of zinc*).

4. They are mostly *readily fusible*, the temperature of fusion varying from −40° C. in the case of *mercury*, to the intense heat of the oxy-hydrogen blowpipe which is required to melt *platinum*.

5. As a rule metals are *good conductors of heat and electricity*, whilst non-metals are bad conductors; but here again they vary much in degree, for carbon is a good conductor.

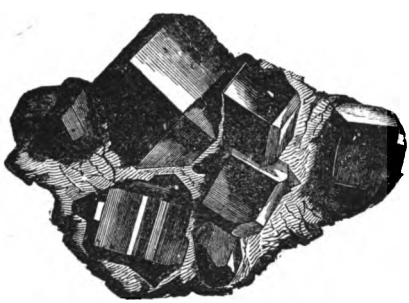

Fig. 91.—Crystals of tinstone (*dioxide of tin*).

6. The metals combine with oxygen to form, as a rule, *basic oxides*, such as *lime*, *alumina*, *soda*, *potash*, etc.; whilst the non-metallic elements, with oxygen, form *acid-forming oxides*, such as *carbon dioxide*, *sulphur dioxide*, and *sulphur trioxide*

7. The *solution of a metal* in a liquid is *attended by a chemical change*; whereas non-metallic bodies, such as *sulphur*, *iodine*, *bromine*, *phosphorus*, etc., may be dissolved in suitable liquids without chemical change.

The **most abundant metals** in the rocks of the earth

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are *aluminium*, *calcium*, *sodium*, *magnesium*, and *potassium*, none of which are found free or in the elementary condition. In addition to these, *iron*, *lead*, *copper*, *mercury*, *zinc*, *tin*, *nickel*, *bismuth*, *antimony*, *gold*, and *silver* are of great **industrial importance.**

These metals all more or less readily combine with oxygen to form **oxides**, with chlorine to form **chlorides**, and with sulphur to form **sulphides.** The oxides will again unite with acids to form **salts.** It is, therefore, as oxides, or sulphides, and chlorides, and other salts, that these metals are met with in the earth.

ALLOYS form a very important class of metallic bodies, for they are largely employed in the arts and manufactures. These substances may be regarded as very **intimate mixtures of two or more metals produced by fusion.** When two metals with widely different properties are dissolved one into the other a new body is yielded with very distinctive characteristics. For example, the alloy "**plumber's solder**" is produced from lead and tin, yet it melts at a lower temperature than either of those metals. This property of **low fusibility** is a special feature of many alloys. **Wood's metal,** an alloy composed of 15 parts of bismuth, 8 parts of lead, 4 parts of tin, and 3 parts of cadmium, melts at just above 60° C., or far below the boiling point of water. By varying the proportions of the components different fusing points are obtained. This principle has been applied in the construction of automatic fire extinguishers and in the manufacture of safety plugs for boiler purposes. Water pipes extending along the ceilings of a building are provided, at short distances

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apart, with plugs of some fusible alloy. In the case of fire, when the heat becomes sufficiently intense, the plugs melt and an automatic flow of water results.

Pewter is an alloy of tin and lead; **brass** of copper and zinc; **German silver** of brass and nickel; **bronze** of copper, tin, and zinc. **Gun-metal, bell-metal,** and **speculum metal** all consist of copper and tin in different proportions. Speculum metal, which is employed for the reflectors of telescopes, has relatively more tin; gun-metal has the least. **Type-metal** is an alloy of antimony and lead, which possesses the peculiar property of expanding at the instant of solidification; it therefore yields a sharp clean casting.

Experiment 157.—Prepare a mixture of 2 parts of *bismuth*, 1 of *tin*, and 1 of *lead*. Introduce it into a small clay crucible placed on a pipeclay triangle supported upon a tripod stand. Melt the mixture at a gentle heat by means of a Bunsen flame. When quite fused pour the liquid out on to a cold stone slab. Break off a small piece of the alloy and place it in some water in a test-tube and boil the water.

Observe that the alloy thus prepared is very like "plumber's solder" in appearance. It melts in the boiling water; its fusion point must, therefore, be below 100° C. If a thermometer is placed in the water so that its bulb is in contact with the alloy, we shall notice that the substance melts at about 94° C.

We learn from this that a **fusible metal** with a very low melting point may be prepared by making an alloy of lead, tin, and bismuth. If we replaced half the tin by an equal weight of cadmium the alloy is **Wood's metal,** which begins to melt at about 60° C.

AMALGAMS.—When mercury is one of the ingredients of an alloy the body is known as an amalgam. Mirrors are "silvered" with an amalgam which contains tin, for this purpose sheets of tinfoil are spread upon smooth surfaces and covered with a layer of mercury; the glass is then pressed thereon.

Various amalgams containing mercury, with silver, and tin are largely employed in dentistry.

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Silver or gold also forms amalgams with mercury which are largely used in electro-plating.

In consequence of the ease with which certain amalgams are formed when mercury comes in contact with clean metallic surfaces, the liquid metal should never be allowed to touch silver or gold articles.

The following is a list of the more **common metallic elements,** and the names of those treated of in this little work are printed in **thick type :—**

COMMON METALLIC ELEMENTS.			
Aluminium.	**Iron.**	Antimony.	Manganese.
Sodium.	**Copper.**	Bismuth.	Chromium.
Potassium.	**Lead.**	**Mercury.**	
Calcium.	**Tin.**	Silver.	
Barium.	**Zinc.**	Gold.	
Magnesium.	Nickel.	Platinum.	

Since the non-metallic elements are very abundant in nature, the earlier pages of this work have been devoted to the consideration of the chief properties of the more important of them, and to the study of the common compounds which they form. Those of the non-metals which we have studied are printed in **thick type** in the following list, and the names of those which are only met with in small quantities are in ordinary type, whilst those which are rare in occurrence are in *italics.*

NON-METALLIC ELEMENTS.			
Hydrogen.	**Oxygen.**	**Nitrogen.**	**Carbon.**
Chlorine.	**Sulphur.**	Phosphorus.	Silicon.
Bromine.	*Selenium.*	Arsenic.	
Iodine.	*Tellurium.*	Boron.	
Fluorine.			

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SUMMARY TO CHAPTER XV.

ELEMENTS - Bodies which cannot be reduced to a simpler form, that is substances from which no other kind of matter can be obtained by any known method. About **seventy** known. Divided into two classes—metals and non-metals.

NON-METALS - About **fifteen** known. One—*bromine*—exists in the liquid state; five—*oxygen*, *nitrogen*, *hydrogen*, *chlorine*, and *fluorine*—are gases; the rest are solids at ordinary temperatures. When solid they are brittle and bad conductors of heat and electricity. They form oxides which dissolve in water more or less readily to *form acids*.

METALS - - About **fifty-five** are known, all of which are solids save mercury. They are all more or less *ductile*, *malleable*, *good conductors of heat and electricity*, and usually have a *high specific gravity*. They combine with oxygen, sulphur, and chlorine to form oxides, sulphides, and chlorides. The oxides of metals are all more or less marked by *basic* in properties. Those oxides of metals which dissolve in water form *alkalies*; they are known as metallic *hydrates*, and we have examples of such in sodic, potassic and calcic hydrates.

ALLOYS - - Intimate mixtures of two or more metals produced by fusion. Most important alloys—*Pewter*, *brass*, *gun-metal*, *speculum metal*, *bell-metal*, *type-metal*, *Wood's metal*, *plumber's solder*, and *German silver*.

AMALGAMS - Alloys which contain *mercury*. These bodies are employed in making mirrors, in electro-p'ating, in dentistry, etc.

QUESTIONS ON CHAPTER XV.

1. What are the chief properties by which you would distinguish a metal from a non-metal?

2. Give a list of the most abundant metallic elements found in the earth's crust.

3. In what condition are metals usually met with in rock?

4. What are the characteristic chemical properties of metals?

5. What is an alloy? How are these bodies produced?

6. Name three common alloys, and give the constituents of each.

7. What is fusible metal, and to what purposes has it been applied?

8. How would you prepare fusible metal? and demonstrate its chief properties.

9. Name the constituents of type-metal, brass, and pewter.

10. What are amalgams? Give examples of the uses of these bodies.

11. Write out the names of all the more common metallic elements, and indicate the most important.

12. Give the names of the non-metallic elements.

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Preparation for Sixteenth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Metallic lead in sheet, wire, rod, and pipe form, specimens of galena, carbonate of lead, white lead, red lead, chrome yellow, litharge and sugar of lead, plaster of Paris, teacup, test-tubes, balance, charcoal, blowpipe, apparatus as shown in Fig. 94, chlorate of potash, black oxide of manganese, nitric acid, hydrochloric acid, sulphuric acid, acetic acid, sulphuretted hydrogen apparatus, and four 8-oz. bottles provided with corks.

EXPERIMENTS TO BE PERFORMED.

"Piece of lead to scrape, and show it is then bright and has 'metallic' appearance.

Show by balance that compared with water it is bulk for bulk much heavier.

Show the metal beaten out into thin sheet; also as wire.

Melt lead in an iron spoon, and cast in a mould.

Show formation of oxide by blowing air on to the melted metal.

Convert the oxide again into metal by strongly heating an intimate mixture of it and charcoal powder.

Heat the oxide to show its further oxidation and the formation of red lead. This oxidation is hastened by adding a little chlorate of potash.

Show the action of nitric acid on lead; also on lead oxide, and the formation of lead nitrate.

Show this as 'a salt,' and prove that it is soluble in water.

Demonstrate that lead is very slightly soluble, and lead oxide almost insoluble in water.

Show the formation of chloride and sulphate of lead by the addition of the respective acids to a solution of lead nitrate. Collect on filter the salts so formed, wash and dry them."—*Science Department's Syllabus.*

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CHAPTER XVI.

LEAD.

Its Colour; a Fresh Surface, Bright, but soon Tarnishes in the Air—Lead is 11½ Times as Heavy as Water—Can be Beaten or Rolled into Thin Sheet, or Drawn into Wire—Melts at Temperature 633° F.—Can be Cast into a Mould—Its Combination, when liquid, with Oxygen—Formation of Lead Oxide—The Oxide has entirely different Properties from Lead—Removal of the Oxygen, when heated with Carbon, and the Formation of Metallic Lead—Formation of Red Lead by heating the Oxide—Solution of Lead by Nitric Acid, and the formation of Lead Nitrate—Similarly to Potassium Nitrate this is to be termed "a Salt"—Its Solution in Water—Other Salts of Lead, Chloride, Sulphate—Formation of Sulphate and Chloride of Lead; their Insolubility in Cold Water—Galena, or Lead Sulphide, one of the ores from which Lead is obtained.

LEAD is widely distributed in nature; it is very rarely found free in rocks, but usually occurs combined with sulphur as **lead sulphide** or **galena**, of which a crystal is shown in Fig. 92. **Lead** also exists in the earth, as **lead carbonate**. Blacklead, or graphite, is not a variety of the metal lead, nor does it contain any of this body, but it is, as we saw in our lesson on carbon, a form of that non-metallic element.

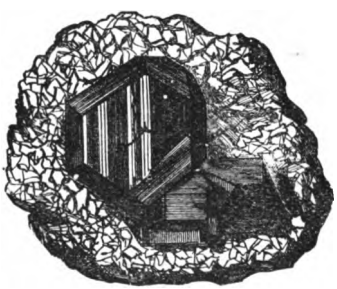

Fig. 92.—Galena (*sulphide of lead*).

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The metal is usually extracted from the sulphide by heating it in air, by which means the sulphur is driven off as sulphur dioxide, and lead remains.

Lead is, bulk for bulk, about 11½ times heavier than water. Its fresh, bright surface has considerable lustre, and a bluish-white appearance; but it soon tarnishes in moist air in consequence of the formation of a thin film of lead oxide. The metal is very flexible and soft, and can easily be beaten or rolled into sheet, or drawn into wire. In consequence of these properties it is largely employed for water-pipes, gas-pipes, roofing purposes, and lining cisterns. Lead melts at a temperature of 334° C.; but when fused with other metals alloys are yielded which melt at much lower temperatures.

Experiment 158.—Scrape the surface of a piece of lead.

Observe that it is bright, bluish-white, with decided metallic lustre. Place it aside so that it is exposed to damp air, and examine after a few days, and notice the film of oxide formed upon the surface.

We learn from this that lead possesses the characteristic properties of a metal and that it will tarnish in air, forming a body known as **suboxide of lead.** Some portion of this is very liable to dissolve in water, especially soft water or rain-water; and, as all soluble salts of lead are poisonous, water which has stood in lead pipes should not be used for drinking.

Experiment 159.—Prepare a mould by mixing in a teacup some plaster of Paris into a pasty mass with water; when nearly set press into the mass a short cylindrical piece of iron, the surface of which should be covered with oil; withdraw the metal, and place the mould aside until dry. Now melt some lead in an iron spoon over a Bunsen flame, and pour the molten metal into the prepared mould and allow it to cool, then withdraw the cylindrical rod of lead which has thus been cast.

Observe that the lead melts, and that its surface becomes quickly coated, whilst in the molten state, with a film of oxide.

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We learn from this that lead melts at a lower temperature than iron, and that the metal may be cast. Molten lead will rapidly combine with oxygen of the air.

Experiment 160.—Select a test-tube into which the cylindrical rod of lead prepared in the last experiment will just fit, and carefully weigh it. Now withdraw the lead, and pour water into the tube until it stands at exactly the same height in the tube as the lead, and weigh again. Deduct, in each case, the weight of the glass.

Observe that the lead is much heavier than an equal bulk of water.

We learn from this that the specific gravity of lead is much higher than that of water; and if in this experiment, after deducting the weight of the glass in in each case from the total weight recorded, we then divide the weight of the water into the weight of the same volume of lead, we obtain $11\frac{1}{2}$, which is very nearly the **specific gravity of lead.**

Experiment 161.—Put a small fragment of lead on a piece of charcoal, and blow the oxidizing or non-luminous flame against it for some time by means of a mouth blowpipe.

Observe the colour of the charcoal in the neighbourhood of the heated metal, and notice that the surface of the lead when cool is covered with a film of a bright coloured oxide.

We learn from this that lead oxide may be formed by blowing air on to the melted metal.

Lead oxide and litharge are the names given to the straw-coloured compound which is produced by blowing a current of air over the surface of strongly heated lead. We saw in our last experiment that the bright, silvery face of the molten metal becomes dull through the action of air. This compound, when strongly heated with sand, yields a glass which is easily fusible. It is also employed in the formation of the glaze for earthenware, and by the assayist as a flux. Lead oxide will dissolve

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in acids to form salts. Metallic lead may be obtained from the oxide by heating it with carbon.

Experiment 162.—Make a mixture of powdered charcoal and litharge, and place it in a hollow in a piece of charcoal. Now heat the mixture strongly in the *reducing flame* of the blowpipe.

Observe that bright metallic globules make their appearance, which are malleable, and will make a dark stain when drawn across paper.

We learn from this that lead oxide may be reduced to the metallic state—*i.e.*, robbed of its oxygen by strongly

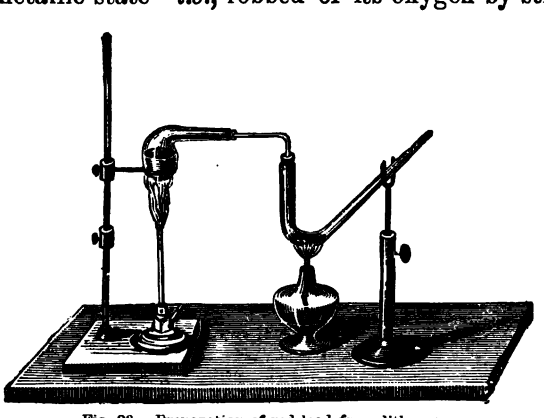

Fig. 93.—Preparation of red lead from litharge.

heating with carbon, in which case the carbon, of course, behaves as a reducing agent.

Red Lead.—Another oxide of lead is known which has a bright red colour, and is termed **red lead.** This body may be prepared from litharge by heating it while exposed to air ; under such conditions the litharge absorbs oxygen and forms a higher oxide of the metal. Red lead is, therefore, richer in oxygen than litharge; this substance is employed in the preparation of soft glass and as a pigment.

Experiment 163.—Introduce a little litharge ; spread out in a thin

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layer into an iron spoon, and heat it to redness. Allow it to remain exposed to the air, in this heated condition, for some time.

Observe that the surface of the litharge slowly passes into **red lead.**

We learn from this that the higher oxide of lead may be prepared by exposing the lower oxide in a heated condition to air.

The preparation of red lead, by the oxidation of litharge, may be hastened by the addition of a little chlorate of potash. By employing a piece of apparatus, as shown in Fig. 93, we may, with ease, produce the body. Oxygen is prepared in the retort on the left, and passes over some litharge which is placed in a piece of hard glass tubing. The litharge is heated by means of a spirit lamp, and thus by the oxygen the change into red lead is quickly brought about.

SALTS OF LEAD are, in many cases, of great commercial importance; they deserve therefore a little attention at this stage. These bodies may be produced by the action of an acid upon metallic lead or the oxide. Lead salts form the basis of many paints. **White paint** is a mixture of carbonate of lead and hydrate of lead, suspended in linseed oil. It is often adulterated with sulphate of barium and zinc oxide. Other lead compounds, such as chromate, chrome yellow, are used for coloured paints.

Lead Nitrate is a white solid salt which may be prepared by dissolving lead or litharge in nitric acid.

Experiment 164.—Place in a test-tube a little lead or litharge, and add nitric acid which has been previously diluted with its own volume of water; warm the mixture.

Observe that red fumes are evolved, and the lead disappears. Pour some of the solution into a dish, and notice that white crystals are deposited which will dissolve in water.

We learn from this that nitrate of lead is a white salt

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soluble in water, which may be obtained by treating lead or lead oxide with nitric acid.

Lead chloride is a salt which is best prepared by adding hydrochloric acid to a solution of lead nitrate, for the acid when cold has very little action upon lead.

Experiment 165.—Place some of the crystals of lead nitrate obtained in experiment 164 in a test-tube and dissolve by shaking with cold water. Now add hydrochloric acid drop by drop.

Observe that a fine white crystalline powder is produced which falls to the bottom of the tube.

We learn from this that lead chloride may be prepared by the action of hydrochloric acid upon lead nitrate, and that it is insoluble in cold water. Water will only dissolve about one-thirteenth of its weight of this substance.

Lead sulphate is a solid white salt which may be most conveniently prepared from a solution of lead nitrate by the action of sulphuric acid. Lead is only very slowly acted upon by dilute sulphuric acid, but it yields more readily to the action of the concentrated acid; for this reason ordinary commercial sulphuric acid which is prepared in leaden chambers always contains some sulphate of lead in solution.

Experiment 166.—Dissolve a few crystals of lead nitrate in distilled water, then add drop by drop dilute sulphuric acid.

Observe that a white precipitate is produced; this should be washed and carefully dried.

We learn from this that **lead sulphate is insoluble** in water, and that it may be produced by the action of sulphuric acid upon lead nitrate.

Experiment 167.—About half fill a clean test-tube with distilled water. Pour into this some ordinary concentrated oil of vitriol, and shake.

Observe that a fine white powder falls through the liquid; this consists of sulphate of lead.

We learn from this that sulphate of lead, which is held

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in solution as an impurity of concentrated oil of vitriol, is thrown out of solution on addition of water.

Lead acetate or **sugar of lead** is a white salt with a somewhat sweetish taste (*it is poisonous*), which is produced by dissolving lead or lead oxide in acetic acid. Other organic acids of meat and fruits, for example citric, will form highly poisonous salts of lead. The action is always far more rapid with the oxide than with the metal. Solder contains lead, and with this the organic acids present in most moist tinned foods may form poisonous compounds. The action is always far more rapid and the danger increased after opening the can, since oxidation is then hastened. *Hence the contents of a can of preserved fruit, vegetables, meat or fish should be taken out directly after opening.*

Lead carbonate is a white insoluble substance which is formed when water containing carbonic acid comes in contact with lead. This body is largely employed under the name of **white lead** as the basis of paints. **Painter's colic** results from lead poisoning, the harmful compounds being introduced frequently through the soiling of food. Epsom salts, sulphate of magnesia, or other soluble sulphate act as **antidotes**, since they convert the lead into **insoluble sulphate of lead.**

Lead sulphide is a black insoluble substance which may be prepared by the action of sulphuretted hydrogen upon lead or any salt of the metal. In consequence of this fact sulphuretted hydrogen is employed largely as a test for lead when even only small quantities are suspected.

Experiment 168.—Pass sulphuretted hydrogen into separate portions of water holding lead nitrate and lead acetate in solution, and lead chloride, lead sulphate, and lead carbonate in suspension.

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Observe that a black precipitate of lead sulphide is produced in each case, which is soon deposited as a sediment in the liquid.

We learn from this that lead may be detected by means of the reagent sulphuretted hydrogen.

Experiment 169.—Select four 6-oz. wide-mouthed bottles. (*a*) Fill one with distilled water which has been boiled to expel all air, and place in this a strip of freshly-scraped lead.

(*b*) Into the second pour some rain water or ordinary distilled water, and also place in this a strip of bright metallic lead.

(*c*) Fill this with ordinary hard water or distilled water to which a few grains of calcium sulphate have been added. Now place in this a strip of bright lead.

(*d*) Into this jar pour rain water or ordinary distilled water, and then add a few grains of litharge.

Cork the bottles and set them aside. After a few days decant a small quantity of the liquid from each, and allow sulphuretted hydrogen to bubble through each portion. This gas may be prepared as directed in experiment 118. Since sulphuretted hydrogen forms with lead salts the black sulphide of lead, the darkening of the liquid on the passage of this gas may be taken to indicate the presence of lead, and the degree of coloration would roughly indicate the comparative quantity of lead dissolved.

Observe that the following results are obtained :—

(*a*) No darkening or coloration of the liquid.

(*b*) The liquid is darkened by the sulphuretted hydrogen.

(*c*) No change in colour.

(*d*) No change in colour.

We learn from this the following facts :—

(*a*) Distilled water has little power of dissolving lead.

(*b*) Rain water, soft water, or distilled water containing air will dissolve lead.

(*c*) Hard water possesses little power of dissolving lead.

(*d*) Lead is dissolved in soft water much more readily than lead oxide.

From this last fact we see how it is that the layer of insoluble oxide and salts formed upon the surface of the lead in pipes and cisterns may serve to protect the users of the water from lead poisoning.

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SUMMARY TO CHAPTER XVI.

LEAD

- **Occurrence in Nature** — It exists only in small quantities free, and is usually found as **galena,** *lead sulphide*; **cerusite,** *lead carbonate.*
- **Preparation** — The metal is chiefly obtained by roasting the sulphide in presence of air, by which means it is converted into lead oxide, which is afterwards reduced by heating with carbon.
- **Properties** —
  - Soft, ductile, malleable. Easily fusible; melting at 334° C. It has a specific gravity of 11·3. It oxidizes in moist air, but not in dry air.
  - *Litharge*, yellow oxide, may be converted into *red lead* by heating and exposing to air.
  - Combines with acids to form salts.
  - Soluble salts of lead are all poisonous.
- **Salts** —
  - Prepared by the action of acids upon metallic lead.
  - *Lead nitrate*, prepared by dissolving lead in nitric acid, is soluble in water.
  - *Lead chloride*, prepared by the action of hydrochloric acid upon a solution of lead nitrate, insoluble in cold water.
  - *Lead sulphate*, prepared by the action of sulphuric acid upon a solution of lead nitrate, insoluble in water.
  - *Lead sulphide* consists of lead and sulphur. It may be produced by the action of sulphuretted hydrogen upon any soluble salt of lead. It is a black substance insoluble in water. Sulphuretted hydrogen is employed as a test for lead.

QUESTIONS ON CHAPTER XVI.

1. How does lead occur in nature?
2. What are the chief physical properties of lead?
3. What are the chief compounds of lead and oxygen, and how may they be prepared?
4. What is red lead? For what purposes is this body employed?
5. Explain fully how you would prepare red lead.
6. What is sugar of lead, and how may it be prepared?
7. What is white lead, and for what purposes is it employed?
8. How would you prepare lead sulphate and lead chloride?
9. How would you demonstrate the action of soft water upon lead?
10. Explain how you would test for lead in water.

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Preparation for Seventeenth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of hæmatite, magnetite, clay ironstone, slag, cast iron, wrought iron, steel. Hydrochloric acid, filter, evaporating basin, iron wire, iron rust, iron poker, blowpipe, black oxide of iron, apparatus as shown in Fig. 95, granulated zinc, nitric and sulphuric acids, specimens of ferrous and ferric salts, solution of potassic ferrocyanide, test-tubes, beakers, and evaporating dish.

EXPERIMENTS TO BE PERFORMED.

"Dissolve cast iron in hydrochloric acid diluted with equal volume of water, show carbon which remains, filter and evaporate the liquid to show the chloride of iron formed.

Heat iron wire by the blowpipe to show the high temperature required to fuse it.

Iron acted on by air and moisture to show its rusting.

Heat iron oxide in a tube and pass hydrogen over it to show formation of water and metallic iron."—*Science Department's Syllabus.*

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CHAPTER XVII.

IRON.

Not used in a Pure Condition; always obtained united with Carbon—Three kinds of Iron: Wrought Iron, Cast Iron, and Steel—Wrought Iron the Purest, and used if the Body is to be formed by hammering—Cast Iron contains most Carbon—Steel used for cutting Instruments; can be made into a Magnet; can be "annealed"—Solubility of all three forms of Iron in Sulphuric, Nitric, and in Hydrochloric Acids; and the formation of Iron Sulphate, Nitrate, and Chloride—Their Solubility in Water—Melting Point of Iron is at much higher Temperature than that of Lead—Comparison of the weight of Iron with that of Water—Its Colour—The ready Action of Moist Air on it—Formation of Rust—Oxidation by Heating—The Action of Steam on Iron when Red Hot—Oxide of Iron heated with Hydrogen or with Carbon parts with its Oxygen, and Iron is left—Oxide of Iron found in the Earth—Hæmatite—Clay Ironstone: a Carbonate of Iron mixed with Clay used as a source of Iron—By heating the Ore Oxide of Iron is obtained—Removal of the Oxygen by heating it to a very high Temperature with Carbon—Formation of Slag from Clay and Lime.

IRON, as we have seen, is one of the most abundant elements in the earth, and it is the most useful and important of all the metals. It is nearly eight times as heavy as water, bulk for bulk. Iron is present in various ores, and it is met with free in small quantities in meteorites, but for commercial purposes it is chiefly obtained from **hæmatite, magnetite,** and **clay ironstone.**

Hæmatite is an oxide of iron which occurs in nodular masses (see Fig. 94). From this fact it is frequently called *kidney ore.*

Magnetite, or magnetic oxide of iron, is the form of ore which yields natural magnets or *loadstones.*

Clay ironstone is the chief iron ore employed in this country. It is an impure carbonate of iron, which is found associated with clay and lime.

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Iron is obtained from its ores by heating them with coal and broken limestone in a blast furnace. The heat decomposes the ore, the carbon of the coal combines with oxygen, and the lime unites with the silica and alumina and forms a glassy **slag.** The lighter slag rises to the surface of the molten mass in the furnace and is drawn off.

Fig. 94.—Hæmatite (*oxide of iron*).

The melted iron falls to the bottom, collects, and is allowed to run into small trenches of sand. The metal thus obtained is not *chemically* pure iron, for it contains varying percentages of carbon, silica, and other substances. As ordinarily classified there are three well-known kinds of iron: viz., cast iron, wrought iron, and steel.

Cast iron is obtained from the ore by smelting, and from this wrought iron and steel are made. Pure iron melts at nearly 1800° C. All the varieties of common iron have lower melting points and contain varying percentages of iron, as shown in the following list:—

Cast iron contains about 90 per cent. of iron, 2 to 5 per cent. of carbon, and melts at 1200° C.

Steel contains about 99 per cent. of iron, ·5 to 3 per cent. of carbon, and melts at 1400° C.

Wrought iron contains about 99·8 per cent. of iron, trace of carbon, and melts at 1500° C.

Cast iron, or pig iron, as it is sometimes called, is reduced from a mixture of ore, limestone, and coal in large

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blast furnaces, which are often as much as one hundred feet in height with a capacity of 10,000 cubic feet. This form of iron contains, as shown above, from two to five parts in a hundred of carbon, which is derived from the coal employed in smelting. Slight variations in the proportion of carbon and other impurities present greatly modify the properties of the metal. Cast iron cannot be welded, as it fuses without previous softening; therefore, on account of its comparatively low melting point, it is best adapted for making castings.

Experiment 170.—Place some small fragments of cast iron in a test-tube, and cover them with hydrochloric acid which has been diluted with an equal volume of water. Warm the mixture, and after the iron has all dissolved pour the whole into a filter. Wash the residue on the filter paper with water. Evaporate some of the filtrate to dryness in an evaporating basin.

Observe that the iron dissolves in hydrochloric acid, that carbon remains on the filter paper, and that a reddish-brown solid, chloride of iron, is obtained on evaporating the solution.

We learn from this that cast iron is soluble in hydrochloric acid, that it contains carbon, and that it forms a soluble salt, **chloride of iron,** with hydrochloric acid.

Wrought iron or malleable iron is the purest common form of iron, and it is this which is used if the body has to be *wrought* into any shape by hammering. It is very fibrous, ductile, and tough, and is easily welded. This form of iron is prepared from pig iron by burning out some of the carbon which is present in that body. The impurities are burnt out in an open reverberatory furnace, the cast iron is melted, and whilst the hot gases are playing over its surface it is kept in constant motion by means of puddling rods. This process is known as **puddling.**

Steel is the hardest variety of iron, and it is consequently

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employed for making armour-plates, cutting instruments, etc. Steel contains less carbon than cast iron, but more than wrought iron, and it may be prepared by fusing these two forms together in suitable proportions. Or it may be made by **carburizing** wrought iron, in which case iron and carbon are packed together and heated for days without melting. The finest kinds of steel are made in this way, but they are very expensive. A cheaper variety known as Bessemer steel is prepared by melting cast iron and blowing hot air through it. When it has been reduced to about the composition of wrought iron a proper proportion of a white variety of pig iron, *speigeleisen*, is added, and the whole is rapidly converted into Bessemer steel.

Steel loses its characteristic hardness and elasticity when heated and allowed to cool slowly, and becomes very like wrought iron. If heated to redness and plunged into cold water it becomes hard and brittle. This process is called **tempering**. By varying the temperature to which it is raised and the conditions of cooling, different degrees of hardness are obtained. This is called **annealing**. Steel may be made into a **magnet** by stroking it with another magnet, or by passing an electric current round it.

Oxides of iron.—When iron is exposed to moist air it loses its bright metallic lustre and becomes "**rusty.**" The rust is an oxide of iron which has been formed upon its surface by the action of the air; the metal will not rust in perfectly dry air, but retains its lustre indefinitely. When iron is strongly heated in air, it burns to form another and higher oxide which contains more oxygen than rust.

Experiment 171.—Heat a piece of iron wire in the Bunsen flame.

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Observe that it becomes red hot, and if the heat is increased by employing the blowpipe the iron burns, and sparks of a black oxide fall off.

We learn from this that the melting point of iron is very much higher than that of lead, and that iron forms when heated **a black oxide.** This oxide of iron was produced in experiments 69 and 85, in the former by burning iron wire in oxygen, in the latter by the action of steam upon iron filings. The same body may be observed in the

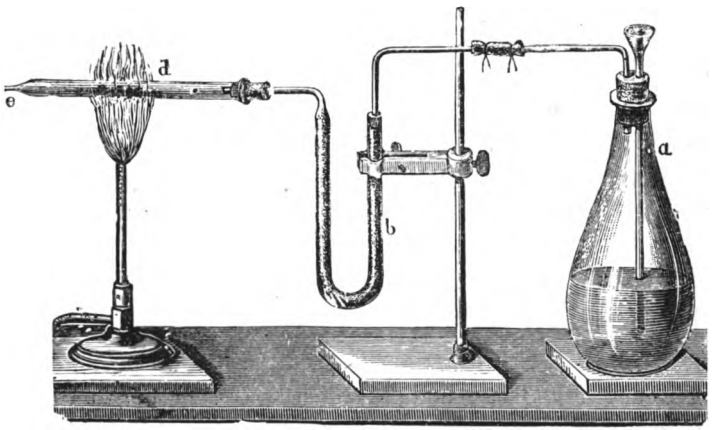

Fig. 95.—Reduction of oxide of iron by hydrogen.

form of thin black scales upon a poker which has been made red hot and then allowed to cool in the air.

Pure iron may be obtained by passing a current of coal gas or hydrogen over oxide of iron heated to redness. In the preparation of the common varieties of iron—all of which we have seen are impure—carbon is employed as the reducing agent. In the following experiment the reduction of the oxide is brought about by the agency of hydrogen.

Experiment 172.—Fit up an apparatus as shown in Fig. 95. On the right the hydrogen generating flask *a*, the delivery tube of which is

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connected with a U tube containing pumice stone saturated with sulphuric acid to dry the escaping hydrogen. The tube *d* contains *pure iron oxide*, which is heated by means of the Bunsen flame. After the hydrogen gas has been escaping for some minutes, and when all the air has been carried out of the apparatus, the oxide of iron should be strongly heated.

Observe that steam escapes from the tube at *e*, and that a fine black powder remains of pure metallic iron.

We learn from this that pure iron may be prepared from its oxide by the reducing action of hydrogen. This finely divided iron is known as *Ferrum Reductum.* When the tube is cold, if a little of the powder is shaken out the particles will take fire as they fall through the air, for they are in such a fine state of division and oxidize so rapidly that they become white hot.

Salts of iron.—All the above-mentioned forms of iron will dissolve in the common acids, viz., hydrochloric, nitric, and sulphuric to form the salts, known as iron chloride, nitrate, or sulphate. There are two series of these bodies known as **ferrous** and **ferric** salts. The former class embraces all those containing the larger proportion of iron to acid ; and the latter includes those which contain the larger quantity of acid. These salts are greenish and brownish red in colour. Bottle glass owes its green colour to **ferrous silicate ;** whilst red clays, red sandstones, and blood, all owe their red colour to the presence of **ferric salts.**

Iron salts may be tested for by means of a solution of ferrocyanide of potassium.

Experiment 173.—Dissolve a small portion of the chloride of iron obtained in experiment 170 in distilled water ; to the solution add drop by drop a solution of ferrocyanide of potassium (yellow prussiate of potash).

Observe that a blue compound is formed by the union of these bodies.

We learn from this that ferrocyanide of potassium is a test by means of which we may identify iron in a solution.

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SUMMARY TO CHAPTER XVII.

IRON

- **Occurrence in Nature** — Iron rarely occurs free; it is usually met with combined with oxygen as *hæmatite* and *magnetite*, as carbonate in *clay ironstone*, and as sulphide in *iron pyrites*.
- **Properties** — Greyish-white metal, capable of taking a high polish. It is extremely tenacious and tough, and very *infusible*; but at a red heat it softens, and can be forged and hammered into any shape. At a white heat it becomes pasty, and, if hammered together, two clean surfaces adhere and become perfectly *welded*.
- **Common Varieties** — *Cast iron* contains carbon; is usually crystalline and brittle. *Steel* contains less carbon; is hard and brittle or very elastic. *Wrought iron* contains least carbon; is malleable and tough. All these varieties are obtained by reducing the iron ore with carbon.
- **Pure Iron** — Produced by passing over red-hot, dry *iron oxide* a current of pure, dry hydrogen.
- **Salts of Iron** — *Prepared* by dissolving the metal or oxide in acids. *Ferrous* and *ferric*; the latter contains most acid; the former are greenish, the latter yellowish red. *Tested for* by ferrocyanide of potassium, which yields a blue precipitate.

QUESTIONS ON CHAPTER XVII.

1. How does iron occur in nature?
2. What are the common forms of iron, and why are they considered impure?
3. Point out the chief differences between wrought iron, cast iron, and steel.
4. Of the common forms of iron which is the purest, and how is it prepared?
5. How would you magnetize iron?
6. How would you prove that cast iron contains carbon?
7. Describe fully how you would prepare a specimen of pure iron.
8. What is "iron rust," and how could you obtain iron from it?
9. Describe fully the action of steam upon heated iron.
10. How would you test for a salt of iron held in solution in water?

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Preparation for Eighteenth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of copper ores, copper pyrites, malachite, cuprite, azurite, copper foil, copper in bar, sheet, and wire, apparatus as shown in Fig. 97, granulated zinc, sulphuric acid, flask, nitric acid, specimens of copper oxide, nitrate, sulphate, and acetate, and of the following alloys—bronze, brass, German silver, gun-metal, bell-metal, speculum metal, tin, and tinfoil.

EXPERIMENTS TO BE PERFORMED.

Heat piece of sheet copper over flame of Bunsen burner. Show formation of black film. Explain its origin.

Take black oxide of copper and heat in hydrogen gas. Show that metal is again formed and that water is produced.

Show action of nitric and sulphuric acids upon copper.

Show that on placing a knife blade in a solution of copper sulphate metallic copper is deposited on the steel.

Show sample of verdigris and explain how formed.

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CHAPTER XVIII.

COPPER.

Its Colour—Does not rust in Air at ordinary Temperatures—Thin Wire melts in Flame of Bunsen Burner—When heated in Air becomes Black owing to formation of an Oxide—Oxide heated in Hydrogen Gas yields up its Oxygen, Water is formed, and the red-coloured Copper is obtained—Action of Acids on all Copper—With dilute Nitric Acid evolves a Colourless Gas, which turns Red in contact with the Air, and the Metal dissolves, forming a blue Solution of Copper Nitrate—Heated with Sulphuric Acid Copper yields Sulphur Dioxide, the same Gas which is formed when Sulphur burns in Air or Oxygen—The Substance formed when Copper dissolves in Sulphuric Acid is, when crystallized from Water, of a fine Blue Colour known as Copper Sulphate or Blue Vitriol—Action of Vegetable Acids on Copper (Verdigris)—Use of Copper in Alloys—A Penny composed of 95 parts of Copper, 4 parts of Tin, and 1 part of Zinc—Bell-Metal and Gun-Metal contain Copper and Tin.

COPPER is one of the most useful of metals; it is largely employed in manufacture, and is extensively used in the preparation of many alloys. It is distinguished from all other metals by its red colour. It is somewhat heavier than iron, having a specific gravity of 8·8. When heated to bright redness it melts, and at a white heat gives off a vapour which burns with a green flame. Whilst the ordinary Bunsen flame will soften iron wire it will melt copper wire, for this metal fuses at about 1090°C. Copper is very ductile and malleable and very tenacious, and it is one of the best conductors of heat and electricity. Thousands of tons of copper are employed for electric wire, domestic

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utensils, ocean vessels, electro-plating, batteries, and for making dyers' vats and pans. It does not rust in air at ordinary temperatures, but if copper is exposed to moist air it becomes covered with a green rust, and when the metal is heated in air it becomes coated with a film of **black copper oxide.**

Nitric acid dissolves it at once, as we saw in experiment 132, forming nitrate of copper. Hot sulphuric acid also dissolves it with formation of copper sulphate; but it is not dissolved by cold hydrochloric acid. All soluble salts of copper are poisonous.

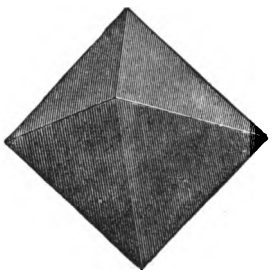

Fig. 96.—Crystal of cuprite (*ruby copper ore, cuprous oxide*).

Copper occurs in nature in large quantities free, but it is chiefly obtained from the ore known as **copper pyrites,** which is a mineral composed of iron, copper, and sulphur, which is found abundantly in Cornwall. It also exists in the minerals **malachite,** which is green, and **azurite,** which is blue, as carbonate, and in **cuprite** as an oxide. Copper is obtained from pyrites by first **roasting** the ore in presence of air; this results in the partial conversion of the sulphide into oxide. Sand is next added, and this combines with the iron to form a fusible slag; thus the iron is removed and the copper obtained in the form of sulphide, or, as it is called, **fine metal.** This is next roasted, and then fused in contact with air, by which means the sulphur is burnt out as sulphur dioxide, and copper remains.

Experiment 171.—Heat a piece of copper-foil or sheet copper in the Bunsen flame.

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Observe the fine film of black oxide which forms upon its surface.

We learn from this that copper oxidizes very readily when heated in presence of air.

Experiment 175.—Hold a piece of fine copper wire in the outer zone of the Bunsen flame.

Observe that it soon melts, and small globules of fused copper fall off; notice that these are coated with a film of the black oxide.

We learn from this that copper melts at a lower temperature than iron, for in experiment 170 we saw that the iron wire was not easily fused.

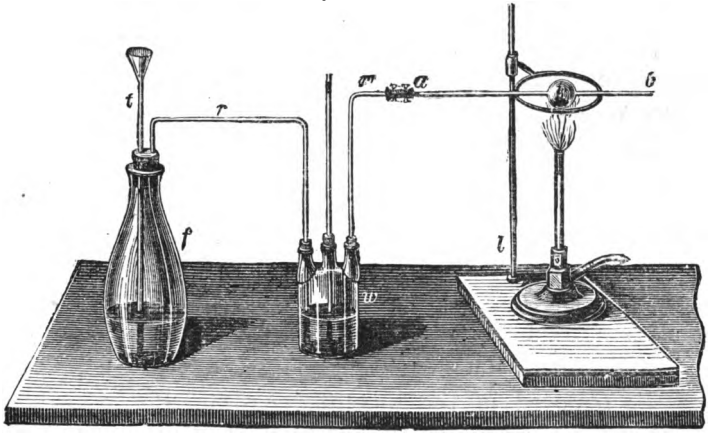

Fig. 97.—Reduction of copper oxide by hydrogen.

Pure copper may be obtained by passing a current of dry hydrogen over heated copper oxide. In such a case the hydrogen combines with the oxygen of the copper oxide and behaves as a reducing agent.

Experiment 176.—Fit up an apparatus, as shown in Fig. 97. Place in the bulb of *a*, *b*, some black oxide of copper; in the flask, *f*, generate hydrogen by the action of zinc upon dilute sulphuric acid; dry the gas by passing it through a wash bottle, *w*, containing strong sulphuric acid. After sufficient hydrogen has been allowed to escape to carry all the air out of the apparatus, heat the bulk, *k*, containing the copper oxide.

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Observe that steam escapes, and that the black oxide of copper slowly changes to red metallic copper.

We learn from this that metallic copper may be obtained from its oxide by the action of heat in the presence of hydrogen; compare the results obtained in this case with those of experiments 86 and 172.

Experiment 177.—Place some copper turnings in a test-tube, add strong sulphuric acid and warm; when the copper has dissolved cautiously add water, and pour some of the solution upon a plate.

Observe that the hot sulphuric acid acting upon the copper causes effervescence, and that the escaping gas has a strong odour; compare experiment 110. The addition of water turns the solution of **copper sulphate blue.** Notice that this yields blue crystals on cooling.

We learn from this that copper sulphate is a fine blue crystalline salt which may be obtained by the action of sulphuric acid upon copper. During the dissolving of the copper the sulphuric acid is so broken up that sulphur dioxide is liberated, which is the gas formed when sulphur burns in air or oxygen. Sulphate of copper is known also as **blue vitriol** or **blue-stone,** it is soluble in cold water, and forms a beautiful blue solution from which the copper may be removed in the metallic state by the action of metallic iron.

Experiment 178.—Prepare a solution of copper sulphate and dip into it a clean knife blade, key, or any piece of iron with a bright surface.

Observe that the iron becomes coated with a deposit of red metallic copper in the form of a film.

We learn from this that iron will displace the metal copper in the solution of copper sulphate; for what really takes place is this—a little of the iron is dissolved to form sulphate of iron, and the copper, liberated by the acid with which the iron combines, is deposited as a bright red layer upon the surface of the iron.

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Copper nitrate is produced by the action of dilute or strong nitric acid upon metallic copper. It is a blue salt which will readily absorb moisture from the air. Although it may be prepared in the crystalline form, when the crystals are exposed to air, they so rapidly absorb moisture that they **deliquesce**.

Experiment 179.—Place in a flask some copper turnings and add dilute nitric acid; gently warm the mixture.

Observe that an invisible gas is given off which turns red in contact with air, and that the metal dissolves, forming a beautiful blue coloured solution of copper nitrate.

We learn from this that copper nitrate may be formed by dissolving copper in nitric acid, and that in its production **colourless nitric oxide** gas is formed, which combines with more oxygen when it comes in contact with the air, forming **red nitric peroxide.** This gas may be collected over warm water by employing a piece of apparatus, as shown in Fig. 47, page 108.

Bright iron dipped into a solution of copper nitrate becomes coated with red metallic copper, as in the case of the sulphate of copper experiment described above. (See number 178.)

Verdigris.—When copper is allowed to remain in contact with organic matter, especially that of vegetable origin, it frequently becomes coated with a bright light green substance known as verdigris. This body is a salt of copper, **copper acetate,** and it is formed by the action of the acid bodies produced, during the decomposition of vegetables, and vinegar upon copper. Since all soluble salts of copper are poisonous, care should be taken not to bring any acid in contact with copper vessels for domestic use.

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Cooking utensils made of copper should always be kept with clean bright surfaces, for they are much more liable to yield to the action of acid bodies when they are coated with tarnish, they then readily form poisonous salts.

ALLOYS OF COPPER are largely employed for industrial and artistic purposes; the more important are:—

Bronze used for English coins contains 95 parts of copper, 4 parts of tin, and 1 of zinc.

Gun-metal, bell-metal, speculum metal all contain copper and tin; the first contains 9 parts of copper and 1 of tin, the second 3 to 6 of copper and 1 of tin.

German silver is an alloy of copper, zinc, and nickel.

Brass is one of the most useful alloys. It is much harder than copper, and is also very ductile and malleable. This alloy is prepared from copper and zinc, and frequently a small quantity of lead. The proportion of copper to zinc is usually 64 parts to 36 parts.

TIN, which is so generally used in the preparation of copper alloys, occurs in the earth as **cassiterite**, tin-stone, which is an oxide of the metal. It is a bright white metal which does not readily tarnish in air, and is used therefore to cover thin plates of copper and iron. **Tinfoil** is an alloy of tin and lead.

ZINC is largely employed in coating wire and sheet iron; that is, in the preparation of what is termed **galvanized iron**. This is done by plunging the iron into melted zinc. The metal occurs in nature chiefly as *sulphide of zinc* in **blende**, and as *oxide of zinc* in **zincite**. Zinc is obtained from its ore by reducing the oxide by heating with carbon and then distilling over the metal.

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SUMMARY TO CHAPTER XVIII.

COPPER

- **Occurrence in Nature**: Copper is found in large quantities free. It also exists largely as *copper pyrites* and *cuprite*, and in smaller quantities it is met with in *malachite*, *azurite*, etc.
- **Properties**: Red, malleable, ductile, tenacious, solid. A very good conductor of heat and electricity. It does not oxidize in air unless strongly heated. It combines with sulphuric acid to form copper sulphate, and with nitric acid to form copper nitrate.
- **Salts**: The more important are :—
  Copper sulphate, blue vitriol, a compound containing copper, sulphur, and oxygen.
  Copper nitrate consists of copper, nitrogen, and oxygen.
  Copper acetate, verdigris, consists of copper, carbon, and oxygen.
- **Alloys**: Copper is largely employed in the preparation of many useful alloys, such as bronze, gun-metal, bell-metal, speculum metal, brass, and German silver.

TIN

- **Occurrence in Nature**: Chiefly as *tinstone* or *cassiterite* in Cornwall, Malacca, and Australia.
- **Uses**: Largely employed in making alloys, of which the most important is *tinfoil*, and for covering copper vessels used in cooking, and for coating sheet iron, which is then known as *tin plate*.

ZINC

- **Occurrence in Nature**: As *blende* or zinc sulphide, *zincite* or zinc oxide, and as *calamine*, carbonate of zinc.
- **Uses**: Employed in preparing galvanized iron. Clean iron is introduced into a bath of melted zinc, when it becomes coated with the alloyed metals.

QUESTIONS ON CHAPTER XVIII.

1. How does copper occur in nature? What is the most important ore of copper found in this country?
2. How is copper extracted from pyrites?
3. How is oxide of copper produced?
4. Explain fully how you would obtain the metal from copper oxide.
5. Describe the action of sulphuric acid upon metallic copper.
6. How would you prepare nitrate of copper?
7. Explain how you would prepare crystals of copper sulphate.
8. What is verdigris? How may it be produced?
9. What is an alloy? Give examples.
10. Name the elements present in brass, bronze, German silver, and gun-metal.

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Preparation for Nineteenth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of mercury, cinnabar, vermilion, red oxide of mercury, nitrate of mercury, mercurous and mercuric chloride, balance, test glasses, nitric acid, test-tubes and glass tubing, tinfoil, and strip of goldfoil.

EXPERIMENTS TO BE PERFORMED.

"Show that to balance a given volume of mercury 13½ volumes of water are necessary.

Boil a little mercury in a test-tube to show that it vaporizes.

Treat mercury with nitric acid and show its solution.

Show that tinfoil is dissolved by mercury, which latter becomes less fluid.

Heat mercuric oxide in a test-tube provided with cork and delivery tube, and collect both the oxygen and the mercury.

Heat mercuric chloride in tube, sealed at one end, with dry sodium carbonate, and show the metallic mercury condensed on the sides of the tube."—*Science Department's Syllabus.*

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CHAPTER XIX.

MERCURY.

A Liquid Metal, but if it be cooled to – 40° F. it is solid—Its Metallic Appearance—Its Weight—Heaviest Liquid known—13·6 times as heavy as Water—Use in the Barometer and Thermometer—Does not Rust or Tarnish in the Air at Ordinary Temperatures—Oxidizes if heated to about 600° F. in the Air, and forms Red Mercuric Oxide—Is readily attacked and dissolved by Nitric Acid—It dissolves many Metals; e.g., *Tin, Lead, etc., Amalgams—Mercury in combination with Sulphur as Cinnabar—Mercury can be obtained from any Salt of Mercury by Heat, Volatilization of Mercury, and the Condensation of the Vapour.*

MERCURY is the only metal which exists in the liquid condition at the ordinary temperatures of the air; if reduced to 40° below the freezing point of water, – 40° C., it passes into the solid state and is then a malleable metal with a crystalline structure very like lead in appearance. The liquid boils at 385° C., and yields a colourless vapour which readily condenses on cooling. Mercury, when pure, retains a bright surface when exposed to air, but if impure, it becomes in a short time covered with a film of tarnish. As explained in Chapter V., when heated to 300° C., and exposed to air, mercury absorbs oxygen, its surface becoming coated with a yellowish red film of mercuric oxide. It is the heaviest liquid known, being about 13½ times heavier than water. This metal is indispensable in the construction

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of thermometers, barometers, the Sprengel air-pump, and other pieces of apparatus employed in physical research. It is also used in silvering and gilding operations. Mercury is attacked and dissolved by nitric acid with the formation of a salt known as nitrate of mercury ; it combines directly with chlorine, forming chloride of mercury, although hydrochloric acid alone will not dissolve it. Hot sulphuric acid

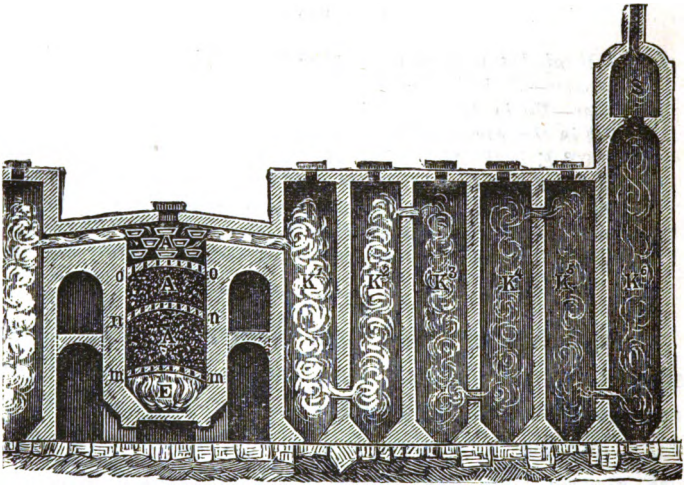

Fig. 98.—Preparation of mercury.

dissolves mercury, forming the sulphate of the metal. When united with sulphur alone it forms mercuric sulphide, a substance which occurs in the earth as **cinnabar** and which yields a valuable pigment known as **vermilion.** The salts of mercury are of considerable value in surgery and medicine, and several furnish useful pigments.

Mercury possesses the property, as explained above, of dissolving many other metals such as tin, lead, gold, silver,

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etc., forming **amalgams**; it is therefore used in the extraction of these metals from their ores. Since mercury can be volatilised by moderate heat it is driven off from any of the above amalgams by heating the body, the mercury being recovered by the condensation of the vapour.

The metal is found native in some parts of the world, but our chief supplies are obtained from **cinnabar.** This mineral consists of sulphur and mercury, and it is obtained from beds or veins, where it occurs associated with native mercury in Peru, Mexico, California, Spain, and Austria. The ore is mixed with lime or iron filings which reduce the **sulphide of mercury**, combining with the sulphur; the mercury is afterwards distilled over into chambers where it condenses, as in Fig. 98, and collected in water.

Experiment 180.—Pour some mercury into a small beaker, and into a second beaker of exactly the same diameter pour distilled water until it stands at the same height as the mercury. Now weigh each carefully, and note down the weights.

Observe that the mercury is very much heavier than water.

We learn from this that the specific gravity of mercury is much greater than that of water. If, in this experiment, we divide the weight of the water into the weight of the equal volume of mercury we obtain 13·6 *as the specific gravity of mercury.*

Experiment 181.—Introduce a few drops of mercury into a test-tube, and heat the lower portion of the tube.

Observe that the liquid vaporizes at once, and is condensed again in the form of minute globules on the upper part of the tube.

We learn from this that mercury can be made to pass into the form of vapour by moderate heat, and that the vapour may be easily re-converted into a liquid.

Mercury does not oxidize at once in air, nor does it

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tarnish when heated even, as we have seen in the above experiment, but when it is kept at a temperature of about 300° C. for some days, exposed to air, it slowly absorbs oxygen, forming **red oxide of mercury.** Nitric acid will oxidize and dissolve the metal at once, forming **nitrate of mercury.** If this salt be heated to about 300° C. ruddy fumes of oxides of nitrogen escape, and a red powder, mercuric oxide or **red precipitate,** remains. This is a much more convenient method of preparing the oxide, and it is therefore usually adopted in preference to the lengthy process described above.

Experiment 182.—Place a few drops of mercury in a test-tube, and then pour on to it strong nitric acid, and heat the mixture.

Observe that the mercury disappears, ruddy fumes escape, and a clear, colourless liquid remains.

We learn from this that mercury is acted upon and dissolved by nitric acid, with formation of nitrate of mercury and liberation of oxides of nitrogen. If some of the solution obtained in this experiment be evaporated to dryness a white crystalline salt may be obtained.

Experiment 183.—Introduce a small quantity of the dry solid salt obtained in the last experiment into a dry test-tube, and heat carefully.

Observe that ruddy fumes escape, and a red powder remains.

We learn from this that by the application of heat nitrate of mercury is broken up into oxides of nitrogen and red oxide of mercury. From this body, we saw in experiment 62, it is possible, by the application of greater heat, to obtain metallic mercury and oxygen.

Experiment 184.—Place some mercuric oxide obtained in experiment 180 in a test-tube provided with a sound cork and delivery tube. Heat the tube and collect the gas, which is given off, over water. Test the gas by means of a smouldering cedar splint.

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Observe that the invisible gas is oxygen, and that mercury collects as a film of metallic globules upon the sides of the colder part of the tube.

We learn from this that mercuric oxide may be decomposed into mercury and oxygen by heat.

Metallic mercury may be obtained from other of its salts by heating them with carbonate of soda.

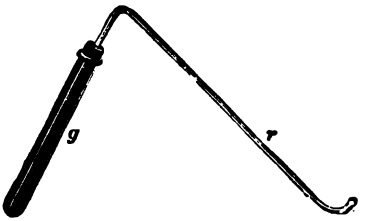

Fig. 99.—Action of heat on red oxide of mercury.

Experiment 185.—Introduce into a long, narrow test-tube a mixture of 1 part by weight of finely powdered **mercuric chloride** and 3 parts of dry soda carbonate, and heat the mixture.

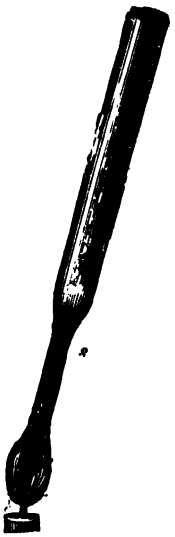

Fig. 100.—Corrosive sublimate vaporizes.

Observe that globules of metallic mercury soon collect upon the cool portion of the tube and form a bright **mirror.**

We learn from this that mercuric chloride is under these circumstances soon decomposed, and that the mercury is liberated from the chlorine. In a similar manner mercury may be separated from any of its salts.

Chloride of mercury when carefully heated alone does not decompose, but vaporizes and sublimes; hence it is known as **corrosive sublimate.** This body is used to a considerable extent as an antiseptic.

Experiment 186.—Introduce into a piece of glass tubing, which has been closed at one end and drawn out to a quill, a small piece of corrosive sublimate. Heat the quill portion of the tube.

Observe the white sublimate which is formed, as shown at *s* in Fig. 100.

We learn from this that corrosive sublimate will sublime without decomposition.

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Another compound of mercury and chlorine is known as **calomel.** This body contains twice as much mercury in proportion to chlorine as **corrosive sublimate.** To distinguish these substances chemists give the former the name **mercurous chloride** (calomel), and the latter they term **mercuric chloride** (corrosive sublimate).

Amalgams.—We saw in Chapter IV. that many solid bodies will disappear in liquids: thus, salt, sugar, and alum all dissolved in water. Many metals in a similar way dissolve in mercury, forming amalgams. The proportion of the solid may be varied so as to make the amalgam thin and liquid, or pasty and solid. Metals which have been dissolved in this way by mercury are said to be **amalgamated.** The solid metal may in most cases be recovered by driving off the mercury by heat, just as salts were obtained from their solutions by evaporation in our earlier experiments. Mercury forms an amalgam so energetically with sodium and potassium, that when these bodies are brought in contact heat and light are evolved.

Experiment 187.—Make a small shallow dish by shaping a piece of tinfoil to the outside of a small crucible. Fit this dish into the mouth of a short wide test-tube, and pour into it a little mercury.

Observe that the mercury soon dissolves its way through the tinfoil, and carries through the hole which it makes the tin which it holds in solution.

We learn from this that tin will dissolve in mercury, forming **tin amalgam.** In a similar manner other amalgams may be prepared.

If a thin amalgam cover the surface of a gold ring or coin the mercury may be removed by nitric acid, for the gold is not acted upon by this body. Heat will also cause the mercury quickly to evaporate from the gold.

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SUMMARY TO CHAPTER XIX.

MERCURY

- **Occurrence in Nature** — Sometimes found native, but usually obtained from *cinnabar*, sulphide of mercury. The metal is obtained by first cautiously oxidizing the ore, and thus converting the sulphur into sulphur dioxide; afterwards the oxide of mercury is reduced.
- **Properties** — Heavy liquid, 13·6 times heavier than same bulk of water. Boils at 385° C., and passes into the solid state at — 40° C. Does not tarnish when exposed to air, and only oxidizes after prolonged exposure to air when heated to about 300° C. Dissolved by hot nitric and sulphuric acids with formation of salts. Not dissolved by hydrochloric acid, but combines directly with chlorine. Mercury dissolves metals forming *amalgams*, from which it may be recovered by heating. It may be separated from any of its compounds by heating with carbonate of soda.
- **Uses** - - Largely employed in the manufacture of barometers and thermometers, in electro-plating, and in making mirrors. It is employed in the extraction of gold and silver, and many of its salts are of high medicinal value.
- **Salts** - - With acids the metal or its oxide will form many salts. The most important are those employed for medicinal purposes and as pigments. The salts of mercury are divided into two classes, viz., the *mercurous* and the *mercuric*. The former all contain more of the metal in proportion to the acid than the latter. For example, *mercurous chloride* contains twice as much mercury as *mercuric chloride*.

QUESTIONS ON CHAPTER XIX.

1. How does mercury occur in nature? How is it extracted from its ores?
2. What are the chief properties by which you would know the element?
3. Describe fully an experiment by which you would demonstrate the action of heat upon mercury.
4. How would you prepare nitrate of mercury?
5. Give two methods of preparing oxide of mercury.
6. How would you obtain mercury from mercuric oxide?
7. What is an amalgam? Name two, and give their components.
8. Describe how you would prepare an amalgam.
9. What practical use is made of the property of amalgamation?
10. What are *calomel* and *corrosive sublimate*?

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Preparation for Twentieth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of common salt, rock salt, sodic carbonate, washing soda crystals, caustic soda, sodic bicarbonate, borax, Chili nitre, metallic sodium, sodic sulphate. Gas jar, test-tubes, pneumatic trough, wire gauze, litmus solution, two flasks, lime-water, and sulphuric acid.

EXPERIMENTS TO BE PERFORMED.

"Recall experiment showing that chlorine is constituent of common salt.

Show that washing soda and sodium bicarbonate evolve carbon dioxide on treatment with an acid; these bodies are compounds of carbon dioxide and sodium.

Show that sodium can be cut with a knife, and that the two portions can be pressed together again so as to form one piece.

Exhibit metallic lustre of sodium; show that it quickly tarnishes in the air.

Show that sodium is lighter than water and decomposes that liquid with evolution of gas (hydrogen).

Collect hydrogen from water by thrusting small piece of sodium beneath test-tube filled with water and standing in basin of water."—*Science Department's Syllabus.*

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CHAPTER XX.

SODIUM.

Common Salt contains a Metal known as Sodium combined with Chlorine—100 Parts of Common Salt contain 39·3 Parts of Sodium, and 60·7 Parts of Chlorine—Carbonate of Soda (washing Soda) contains Sodium—Sodium obtained by strongly heating Carbonate of Soda with Charcoal—Sodium one of the lightest Solids known—Swims on the surface of Water and decomposes that Liquid with evolution of Hydrogen and formation of the Alkali Soda—Other Properties of the Metal Sodium; its low fusibility and softness; its tarnishing in Air—Preservation of Sodium from the action of Air by being kept in some Liquid lighter than Water and free from Oxygen.

COMMON SALT is one of the most abundant minerals in the earth. It is so very soluble in water that it is found very widely distributed in that body. Sea water contains about 2·5 per cent. of salt, and the waters of some salt lakes contain as much as 30 per cent. Common salt also occurs in all fertile soil, and is distributed generally through rock in small quantities: sometimes it occurs in beds nearly pure; it is then known as **rock salt.** We saw in experiment 129, that when the **metal sodium** was burnt in chlorine, common salt was formed. In thus forming the compound 39·3 parts by weight of sodium always combine with 60·7 parts of chlorine to form 100 parts of common salt.

SODIUM does not exist free in nature, and although it may be isolated from its compounds by chemical means, yet

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it has such a powerful affinity for other elements that it is only with the greatest difficulty kept for any length of time in the free or uncombined state. In nature it is also found combined with carbonic acid as **soda carbonate**, with boron and oxygen in **borax**, and with nitric acid in **Chili saltpetre** or **Chili nitre.**

Sodium is obtained in the metallic state from carbonate of soda by strongly heating that body with finely divided carbon. The carbon combines with the oxygen and escapes in the gaseous form, whilst the sodium distils over at the high temperature and is collected in an atmosphere which contains no oxygen. It is usually condensed under the surface of rock oil. Sodium is one of the lightest known solids; it will float on water, being somewhat lighter than that body. It is a soft metal, which may be easily moulded with the finger and cut with a knife; the cut pieces may readily be welded by pressing them together. The freshly cut surface has a bright silvery-white lustre, which, however, is soon destroyed by exposure to air, for it rapidly oxidizes, and a thin crust is formed of **oxide of sodium**. So great is the chemical affinity of sodium for oxygen that it will decompose water combining with its oxygen and liberating hydrogen. Since both air and water act upon this element, and as it possesses the power of abstracting oxygen from compounds containing it, the metal must be preserved either in a gas which contains no oxygen, or in a liquid which is lighter than water, and which contains no oxygen. Naphtha is usually employed for this purpose.

Experiment 188.—Cut a piece of sodium about the size of a small pea and cast it into water which has been coloured with red litmus solution.

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Observe that the sodium fuses and floats upon the surface of the water in globular form, its specific gravity being ·973 ; and that the reddened litmus is turned quite blue.

We learn from this that sodium is lighter than water, it has a low melting point, and that by its action upon water it produces an alkali. If some of the liquid be evaporated to dryness a white solid body is obtained which is known as **sodic hydrate** or **caustic soda**. The sodium combines first with oxygen of the water to form sodium oxide, this then unites with water to form sodic hydrate. Caustic soda is a very powerful alkali, and it is for this reason largely used in making soap.

Experiment 189.—Cut a piece of sodium and press the cut fragments together.

Observe that sodium can easily be cut with a knife or moulded with the fingers, that it may be welded by simple pressure, and that the freshly-cut surfaces become coated with a white film.

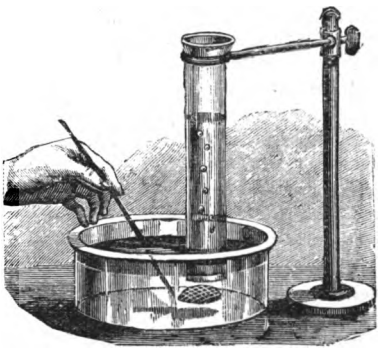

Fig. 101.—Hydrogen liberated by action of sodium on water.

We learn from this that the metal sodium is soft and malleable, and that it rapidly tarnishes in air.

Experiment 190.—Fold a *small fragment* of sodium in a piece of *fine wire gauze*, and by means of a pair of crucible tongs quickly introduce it under the mouth of a gas jar containing water supported in a trough

Observe that the sodium acts upon the water, and that an invisible gas rises through the tube and displaces the water. After a time remove the jar, and notice that the gas is colourless, inodorous, and tasteless, and that it takes fire on the application of a burning match.

We learn from this that hydrogen is liberated by the action of sodium upon water.

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The most important **salts of sodium** are *sodium chloride*, *carbonate*, *nitrate*, and *sulphate*.

Sodium chloride has been proved to be a compound of sodium with chlorine. It occurs widely distributed in nature, and is present in the ashes of all plants and animals; hence it is present in all fertile soils and in the foods taken by animals. This salt is employed in the preparation of both sodium carbonate and hydrochloric acid, for when common salt is acted upon by sulphuric acid, hydrochloric acid is liberated, and sulphate of soda remains. When this latter body is heated with carbon and chalk it yields the sodic carbonate of commerce.

Sodium carbonate occurs in nature in small quantities. In the so-called alkali belt of the western United States it often forms incrustation from an inch to a foot in thickness. The body is prepared in this country from salt by the method given above. It is made in enormous quantities, and its manufacture constitutes one of the greatest chemical industries. This substance is largely employed in soap and glass-making, also in calico-printing and dyeing, as well as for laundry and domestic purposes. Common washing soda is carbonate of soda, and it is therefore composed of the chemical elements carbon, oxygen, and sodium. A compound very similar in composition, but containing some hydrogen, is known as **bicarbonate of soda.** This salt is prepared by passing carbon-dioxide into water holding carbonate of soda in solution.

Experiment 191.—Fit up two flasks with thistle funnels and delivery tubes, and place in one some crystals of washing soda, and in the other a small quantity of bicarbonate of soda. Introduce into each flask a little sulphuric acid by means of the thistle funnels, and allow the escaping gas to pass into lime-water in each case.

Observe that the two bodies effervesce on the addition of the acid, and that the escaping gas in each case turns lime-water milky.

We learn from this that both washing soda and bicarbonate of soda contain carbon-dioxide. We have seen above that they may be regarded as compounds of sodium oxide with carbon dioxide in each case; the former contains more sodium in proportion to carbon and oxygen than the latter, which contains some hydrogen.

From carbonate of soda, by heating with milk of lime, **caustic soda** is prepared for industrial purposes.

Sodium sulphate is a compound containing sodium, sulphur, and

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oxygen. It is prepared by the action of sulphuric acid upon other salts of sodium. It is chiefly used for medicinal purposes, and is known as Glauber's salt. Sodium sulphate occurs as crystals, which readily lose water when exposed to air, and fall to a fine white powder; this property, which is possessed by many other salts, is known as **efflorescence.**

Sodium nitrate occurs as deposits in Chili and Peru; it is chiefly of value for the manufacture of nitric acid, and for manuring purposes.

SUMMARY TO CHAPTER XX.

SODIUM

- **Occurrence in Nature** — Sodium is never met with free in nature, but is found in waters and rocks, chiefly as sodium chloride and sodium nitrate, and in small quantities in the native state as sodium carbonate.
- **Properties** — A soft, easily fusible, silver-white solid, which is lighter than water. It combines with dry oxygen to form sodium oxide, which in the presence of water or aqueous vapour at once forms sodium hydrate or caustic soda. Sodium decomposes water liberating hydrogen.
- **Salts** —
  - *Sodium chloride*, common salt, consists of sodium and chlorine.
  - *Sodium carbonate*, washing soda, consists of sodium, carbon, and oxygen.
  - *Sodium bicarbonate* is made up of sodium, hydrogen, carbon, and oxygen.
  - *Sodium sulphate*, Glauber's salt, composed of sodium, sulphur, and oxygen.
  - *Sodium nitrate*, Chili nitre, consists of sodium, nitrogen, and oxygen.

QUESTIONS ON CHAPTER XX.

1. How is sodium found in the earth?

2. How has metallic sodium been obtained from sodium carbonate?

3. What chemical elements are present in each of the following bodies:—(*a*) common salt, (*b*) washing soda, (*c*) Glauber's salt?

4. What are the chief properties of the metal sodium?

5. What is the action of sodium upon water? Describe fully the properties of the gas which is given off.

6. Point out the industrial uses of the more important salts of sodium.

7. Name the chemical elements which are present in all the more common compounds of sodium.

8. How is caustic soda prepared? To what class of substance does this body belong?

9. What do you understand by the term efflorescent? Name some salt which exhibits this property.

10. What gas is given off when an acid acts upon washing soda?

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Preparation for Twenty-first Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of acetic acid, tartaric acid, oxalic acid, fats, oils, resins, starch, gums, sugar, spirit, gluten, gelatin, vaseline, paraffin, naphtha, and other organic bodies, test-tubes, lime-water, soda lime, litmus, and apparatus as shown in Fig. 103.

EXPERIMENTS TO BE PERFORMED.

Show that on heating any ordinary vegetable or animal substance carbon is left behind.

Show that when an organic body is burnt in air carbon dioxide is produced.

Prove that decomposition of gelatin by heating leads to evolution of ammonia. Show the action of escaping gas upon red litmus.

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CHAPTER XXI.

CARBON COMPOUNDS.—INTRODUCTORY.

Large numbers of Substances are met with in Plants and Animals which are not found in the Earth—Most of these Bodies contain Carbon—The other Elements united with the Carbon are Hydrogen, Oxygen, Nitrogen—Some Bodies are composed of all these Elements; others of only two of them—Many of these Bodies when heated leave a black residue of Carbon; when this is more strongly heated it burns away—The great Number of these Carbon Compounds and the great Difference in their Properties—Some are Acids—e.g., *Vinegar (Acetic Acid) and Tartaric Acid—Some are Salts*—e.g., *Fats, Tallow, Butter—Some are Neutral Bodies*—e.g., *Sugar, Starch, Spirit.*

TWO BRANCHES OF CHEMISTRY.—The preceding chapters have been devoted to the consideration of the nature and properties of many chemical compounds and elements which are found in the earth and in common objects. All the non-elements which have been studied fall under the head of **inorganic compounds** or **mixtures.**

Inorganic Chemistry is the name given to the branch of study which treats especially of the composition of minerals and unorganized materials. But there is another class of bodies which includes an enormous number of substances which contain carbon, all of which are more or less closely related; the number of these known is so great, and is so rapidly growing, that a branch of

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chemical study known as **organic chemistry** is specially devoted to their investigation.

ORGANIC CHEMISTRY was formerly defined as the chemistry of the compounds formed by plants and animals ; but it has a much wider range and may now, rather, be described as the **chemistry of the carbon compounds.** New organic compounds are constantly being discovered, and although many of them are obtainable from plant and animal substances, yet a very large number have been built up by purely **synthetical chemical methods.** In one particular they all agree—viz., that carbon is always the central or grouping element. It was believed that the complex substances found in plants and animals owed their peculiar composition to the action of some **so-called vital force**, which was considered as quite outside the ordinary laws of chemical action ; we know now that this is not the case, but that the same laws of combination and decomposition obtain in the building up and breaking down of organic compounds as in the formation of inorganic bodies. The division of substance into organic and inorganic is merely adopted for convenience, and in consequence of the fact that all members of the former class contain the element carbon, and they possess certain well-marked properties in common. Formerly it was believed that the so-called organic bodies could not be built up from their elements as inorganic bodies could, although it was known that they could be analyzed ; *it was held that they were only the product of living agents.* In the year 1828 **urea**, a substance present in urine, was first artificially prepared, and since then enormous progress has been made in the study

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of organic compounds, and chemists have prepared many thousands of organic bodies by purely synthetical methods. The possibility of preparing quite new organic substances depends upon the fact that carbon differs from most other elements in one very important particular, and that is, it possesses a wonderful power of linking itself with other elements and combining with other elements to form groups, which in their turn behave as elements. These groups, which often play the part of chemical elements in the building up or breaking down of organic compounds, are known as **compound carbon radicles.** Taking advantage of this fact, chemists have been able to prepare many organic products in cheaper and purer forms than those in which nature herself prepares them.

Among common objects the following are a few of the best-known organic compounds: *vinegar*, *tartaric acid*, *fat*, *tallow*, *oil*, *butter*, *sugar*, *starch*, *gum*, *gluten*, and *spirit*. It is well known that these bodies are directly derived from animals or vegetables, but other well-known substances which are often simpler in composition, and which are only indirectly obtained from living things, are included in the great class of organic compounds. For example, *marsh gas*, *coal gas*, *paraffin*, *benzoline*, *vaseline*, *naphtha*, *benzole*, and *carbolic acid* are all organic substances.

Some of the better known organic bodies consist of the **two chemical elements, carbon and hydrogen;** as examples of these we have *marsh gas*, *paraffin*, *turpentine*, and *vaseline*. Others contain the **three chemical elements, carbon, hydrogen, and oxygen.** In this class we have *starch*, *sugar*, *gum*, *fat*, *tallow*, *butter*, *alcohol*, *acetic*

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acid, and *tartaric acid*. Some are composed of four elements, and are distinguished as **nitrogenous compounds** when they contain nitrogen in combination with carbon, hydrogen, and oxygen. Among the best known of the nitrogenous organic bodies we have *gelatin*, *gluten*, and *albumen*. Organic compounds, like inorganic bodies, differ immensely in their chemical characters and in their physical properties and condition. *Vinegar*, *tartaric acid*, *oxalic acid*, are all **acids**. Some are **salts**; for example, *butter*, *fats*, and *oils*; whilst *spirit*, *sugar*, *starch*, and *gluten* are **neutral bodies**.

As before mentioned, all organic compounds contain carbon; and associated with this element are one or more of the following—hydrogen, oxygen, and nitrogen. When such bodies are burnt the carbon combines with oxygen to form carbon dioxide, and the hydrogen unites with oxygen to form water. When an organic body contains a large proportion of carbon, it is not oxidized so rapidly as when hydrogen is largely present, and if an organic substance be heated in only a very limited supply of air, much of the carbon is left behind as a black residue. From this fact we are able, in a large number of cases, by a very simple experiment, to determine whether a body contains carbon or not, and thus decide whether it is organic or inorganic.

Experiment 192.—Introduce into a clean, dry test-tube a piece of starch about the size of a small pea. Hold the tube in a horizontal position and heat the end of it in a Bunsen flame till it is red hot.

Observe the drops of water which collect in the cool portion of the tube, and the black residue of carbon which remains.

We learn from this that carbon is present in starch; and, if the starch was dry, the production of water indicates the presence of hydrogen and probable presence of oxygen

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in starch. Since air was not excluded from the tube, we may argue that the oxygen present in the water may have been derived from the air; but further consideration will lead us to the conclusion that such a large quantity of water could hardly have derived its oxygen from such a small bulk of air as the tube contained.

Other black substances are known besides carbon; it is

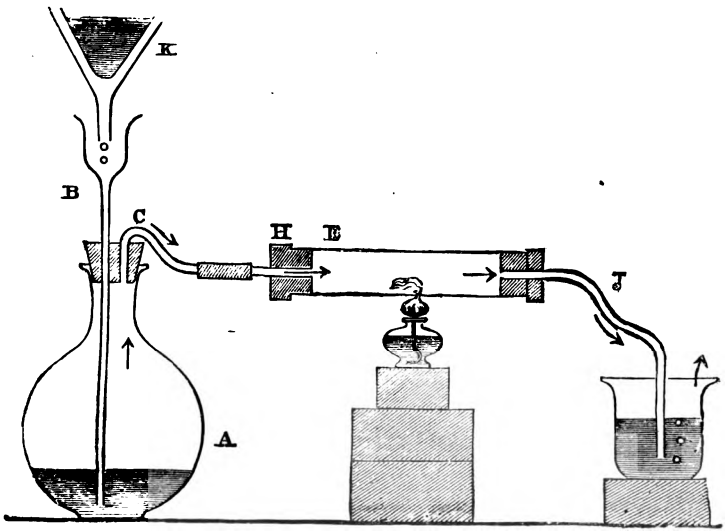

Fig. 102.

best, therefore, in order to prove absolutely that carbon is present in any body to perform the following experiment.

Experiment 193.—Fit up a piece of apparatus as shown in Fig. 102. A is an ordinary wash bottle from which the usual tubes have been removed. B is a thistle tube which passes below the surface of the water in the wash bottle. At C a short delivery tube is fitted bent at right angles. E is a piece of combustion tubing provided with perforated corks at each end. Through H a tube passes which is connected by means of indiarubber tubing with C. Before fixing the corks place a small piece of carbon, obtained in experiment 192, in the tube, and let J dip into a beaker

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containing clear lime-water. K is a large funnel fitted with a filter paper, which is placed just above the thistle tube B. Place the combustion tube on a tripod stand, and heat the whole length first; and then especially the part where the carbon is. Pour some water into K; this will flow slowly into the wash bottle through the thistle tube and drive the air on in the direction of the arrows. Thus air is forced over the heated charcoal, which burns and produces an invisible gas, which is driven forward into the lime-water contained in the beaker on the right.

Observe that the lime-water turns milky and that the milkiness is at once destroyed, and the liquid cleared with effervescence on the addition of a few drops of hydrochloric acid.

We learn from this that when the organic body, starch, is heated it leaves a black residue, and that when this solid substance is more strongly heated in air it burns to form carbonic acid gas. Since carbonic acid gas is produced by burning bodies which contain carbon, we may infer that the black residue which remains after heating starch is carbon.

Test for Nitrogen.—On heating organic bodies containing nitrogen decomposition takes place, and, in most cases, the nitrogen combines with the hydrogen present in the substance to form ammonia, which gas may be detected by its action upon litmus paper, as we saw in experiment 146. The decomposition of the nitrogenous organic body is promoted by the addition of soda-lime.

Experiment 194.—Introduce into a test-tube a few fragments of glue or gelatin, and strongly heat the tube.

Observe that water is condensed on the sides of the tube, and that strong pungent fumes escape which possess a decided ammoniacal odour, and which will turn moistened red litmus paper blue. Notice that a residue of carbon remains, and that the decomposition is promoted by the addition of soda-lime or quicklime.

We learn from this that certain organic bodies when strongly heated give off ammonia; and since we have seen that ammonia consists of the two chemical elements, nitrogen and hydrogen, it follows that such organic bodies must contain nitrogen.

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SUMMARY TO CHAPTER XXI.

INORGANIC CHEMISTRY — The study of minerals and unorganized compounds and mixtures, and the elements of which they are composed.

ORGANIC CHEMISTRY — The *chemistry of the carbon compounds.* These bodies chiefly make up plants and animals, and are largely obtained from the organic kingdom. Many have been built up by purely chemical methods from their elements.

ORGANIC COMPOUNDS —

Organic bodies always contain carbon united with some other element or elements.

Carbon and *hydrogen* make up marsh gas, fire-damp, paraffin, coal gas, turpentine, vaseline, etc.

Carbon, hydrogen, and *oxygen* together make up starch, gum, sugar, fat, tallow, butter, alcohol, acetic acid, tartaric acid, malic acid, oxalic acid, citric acid, etc.

Carbon, hydrogen, oxygen, and *nitrogen* are the components of gelatin, gluten, and white of egg.

Common acids, tartaric acid of grapes, acetic acid of vinegar, malic acid of apples, citric acid of lemons, and oxalic acid of rhubarb.

Common salts, butter, tallow, oils, and fats.

Common neutral bodies, alcohol, gelatin, gluten, starch, and sugar.

QUESTIONS ON CHAPTER XXI.

1. What are the two branches into which the study of chemistry is divided?

2. What are organic compounds? Give examples.

3. Why is organic chemistry sometimes called the chemistry of the carbon radicles?

4. How would you distinguish between an organic and an inorganic compound?

5. Write out the names of twenty common substances, and say whether each is organic or inorganic.

6. What is the action of heat upon organic bodies?

7. What chemical elements are present in each of the following bodies:—*spirit, fat, gelatin, gum,* and *paraffin*?

8. How would you prove that an organic body contained nitrogen?

9. Name some common organic acids.

10. Sketch and describe the apparatus you would employ in seeking for carbon in an organic substance.

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Preparation for Twenty-second Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of vinegar, white vinegar, aromatic vinegar, acetic acid, sodic acetate, iron acetate, lead acetate, copper acetate, soda carbonate, iron filings, lead oxide, ice, dilute sulphuric acid, litmus, apparatus as shown in Figs. 104 and 105, test-tubes, beakers, two evaporating basins, watch glasses, tripod stand, and sand bath.

EXPERIMENTS TO BE PERFORMED.

"Show that vinegar has the properties of an acid, and that a salt is formed on neutralizing it.

Show that iron is acted on and dissolved by acetic acid.

Make sugar of lead by dissolving lead oxide in acetic acid, and crystallize out the salt.

Point out disappearance of the pungent odour of the acid on neutralization by potash or soda.

Demonstrate the liberation of the acid as indicated by the odour on addition of sulphuric acid to sodium acetate, and show that it can be separated from the liquid by distillation."—*Science Department's Syllabus.*

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CHAPTER XXII.

ACETIC ACID.

One form of Dilute Acetic Acid is known as Vinegar—Formation of Acetic Acid when Beer or Wine exposed to the Air becomes Sour—The Spirit present combines with Oxygen of the Air and forms Acetic Acid—The Presence of a kind of Fungus called Mycoderma Aceti necessary to cause this Oxidation—Large amount of Vinegar is made from poor kinds of Wine and Beer—Action of Vinegar on Blue Litmus and on Sodium Carbonate—Vinegar is also made by heating Wood in a Retort; a great many Bodies distil over, among them Acetic Acid—The pure Acid has a very pungent smell, and has all the Properties which are characteristic of the Acids—Boils at 246° F.—Dissolves in Water—It is composed of Carbon, Hydrogen, and Oxygen in the proportion of 40·0 parts of Carbon, 6·7 parts of Hydrogen, and 53·3 parts of Oxygen—It is like Sulphuric Acid neutralized by Alkalies—Iron put into it is slowly dissolved, Hydrogen being given off—Oxide of Lead dissolves, forming a Salt, and if the Solution be evaporated a White Crystalline Body called "Sugar of Lead" is formed which is Lead Acetate—The smell belongs only to the Acid, not to Salts—Sodium Acetate has no smell; add to it Sulphuric Acid and warm, when the smell shows that the Acid has been liberated and that it is Volatile.

ACETIC ACID is the commonest and most important of the organic acids. It is naturally produced in small quantities and in a dilute condition during the souring of wine or other alcoholic liquors. When pure it is a clear colourless liquid with a powerful pungent odour, which boils at 118° C. and freezes at about 17° C. The acid is somewhat heavier than water, having a specific gravity of 1·063. It has a sharp sour taste, will neutralize bases

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forming salts, reddens blue litmus, and causes carbonate of soda to effervesce, and thus possesses all the characteristic properties by which we distinguish an acid.

Vinegar is an impure variety of acetic acid, which is produced by the fermentation of the poorer varieties of wines and beers. The alcohol present in the liquid is, by the action of oxygen, converted into acetic acid. The result is brought about under the influence of a low fungoid form of life known as the *mycoderma aceti*, vinegar plant, or "mother of vinegar." This lowly organism lives, grows, and multiplies in the souring liquid, absorbing oxygen from the air and oxidizing the alcoholic liquid into vinegar. The process is spoken of as **acetous fermentation.** In the "quick process" for making vinegar, weak alcohol is made to trickle through casks containing wood shavings and pierced with holes for the admission of air. The shavings become coated with mycoderma and the oxidation proceeds very rapidly. The strongest vinegar of commerce contains about 5 per cent. of pure acetic acid. From the dilute solution of acetic acid known as vinegar, the pure acid is obtained by first producing a salt by neutralizing the fermented liquid with lime; thus calcium acetate is formed; this is afterwards distilled with sulphuric acid, and yields nearly all the acetic acid of commerce. **Aromatic vinegar** is acetic acid in which fragrant essential oils have been dissolved.

Acetic acid is also produced during the **distillation of wood.** The wood is heated in large iron retorts, closed so as to exclude air, and the vapours which come over are condensed in a receiver; the crude distillate is known as

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pyroligneous acid, which name implies, the acid obtained by heating wood. This product of **destructive distillation** contains tarry substance, wood naphtha, and acetic acid. The liquid portion is separated and redistilled; the more volatile **wood naphtha** comes over first and the **wood vinegar**, which has a higher boiling point, after.

Acetic acid is composed of carbon, hydrogen, and oxygen, in the proportion of 40 parts of carbon, 6·7 parts of hydrogen, and 53·3 parts of oxygen in 100 parts by weight.

Experiment 195.—Pour into two small beakers about 10 cc. of vinegar. To the first add a few drops of blue litmus solution, and to the second add a small crystal of sodic carbonate.

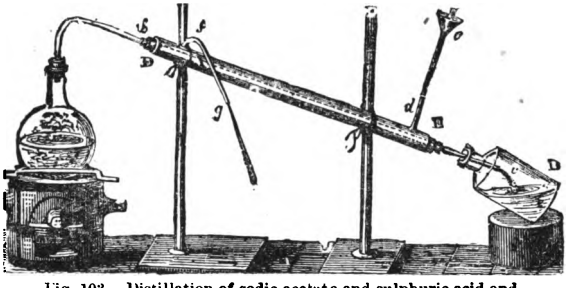

Fig. 103.—Distillation of sodic acetate and sulphuric acid and preparation of acetic acid.

Observe that in the first case the blue litmus is turned red, and in the second the carbonate of soda dissolves with violent effervescence, and the escaping gas turns lime-water milky.

We learn from this that vinegar behaves as an acid.

Experiment 196.—Neutralize a small portion of vinegar by means of solid sodic carbonate, concentrate the liquid by careful evaporation, and allow a portion to cool in a watch-glass or in an evaporating dish.

Observe that crystals are produced of a salt.

We learn from these two experiments that the acetic acid of vinegar will decompose sodic carbonate with liberation of carbon dioxide, and that it forms **sodium acetate.** From this salt it is possible to prepare pure acetic acid by distillation with a stronger acid.

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Experiment 197.—Heat about 50 grains of crystallized sodic acetate in an evaporating dish until the water of crystallization is expelled, then allow the residue to cool. Break it into a coarse powder, and introduce it into a small flask; add an equal bulk of strong sulphuric acid. Allow the flask to stand for some time; then apply heat, and distil over the acid.

Observe that a vapour comes off which condenses to a clear, colourless liquid in the bottle at *c*, as shown in Fig. 103.

We learn from this that sulphuric acid displaces the acid in sodic *acetate* and forms sodic *sulphate*, and as the acetic acid is volatile it may be collected by distillation.

Experiment 198.—Pour some acetic acid into a beaker.

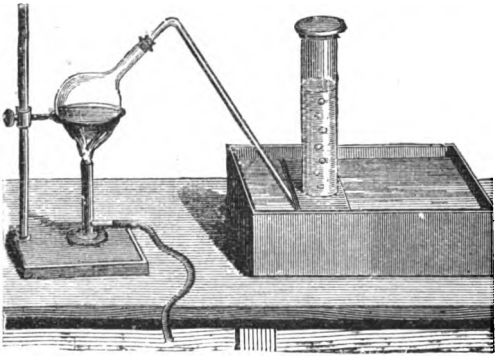

Fig. 104.—Iron dissolved by acetic acid with liberation of hydrogen.

Observe that it is colourless, and has a strong pungent odour, which is increased by warming. Place the beaker in a basin containing a freezing mixture of pounded ice and salt, and notice that the acid becomes solid.

We learn from this the chief properties of acetic acid, including the fact that at low temperature it forms beautiful, colourless, transparent, needle-like and plate-like crystals.

Experiment 199.—Pour a few drops of acetic acid into a tumbler of water, and smell the solution; then taste the liquid.

Observe that the acid at once dissolves in the water, and that the resulting solution has a pleasant sour flavour.

We learn from this that acetic acid is soluble in water. The **white vinegar** of commerce is dilute acetic acid.

ACETATES.—Acetic acid will not only act upon metallic carbonates, as above proved, but it will dissolve some metals and it will combine with many metallic oxides to form salts. The salts produced by the union of acetic acid and a base are all known as **acetates.** The most important

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are: **acetate of soda, acetate of lead, basic acetate of lead** (*Goulard's solution*), **acetate of copper, basic acetate of copper** (*verdigris*), **acetate of alumina,** and **acetate of iron;** the last two are employed largely as mordants in dyeing and calico-printing.

Experiment 200.—Place some scrap iron in a flask, as shown in Fig. 104; then cover it with acetic acid. Provide the flask with a cork and delivery tube; allow the latter to dip into water in a pneumatic trough over which a flask full of water should be inverted. Heat the flask.

Observe that the iron slowly dissolves with effervescence, and that an invisible gas collects in the tube, which on examination proves to be hydrogen. The salt of iron which is produced may be obtained by the evaporation of the clear liquid which remains in the flask.

We learn from this that acetic acid will act upon and dissolve metallic iron; the hydrogen of the acid being liberated, the metal takes its place. The salt which is thus produced is **acetate of iron.**

Experiment 201.—Introduce into an evaporating basin a small quantity of lead oxide and cover it with acetic acid, and warm the mixture.

Observe that the lead oxide dissolves, and that crystals are formed when the solution is allowed to cool.

We learn from this that lead oxide will dissolve in acetic acid with formation of **lead acetate** (*sugar of lead*).

SUMMARY TO CHAPTER XXII.

ACETIC ACID

- **Preparation** — Produced in small quantities by *acetous fermentation* of dilute alcoholic solutions, in which the alcohol is converted into acetic acid through the agency of a fungus, *mycoderma aceti*. The pure acid is prepared by the distillation of sodic, calcic, or potassic acetates with sulphuric acid.
- **Properties** — Pure acetic acid is a colourless liquid with sour taste and pungent odour. It freezes at 32° F. to form a colourless, crystalline mass, and boils at 246° F. It dissolves at once in water in all proportions, and combines with bases to form salts known as *acetates*.
- **Uses** — In its concentrated and dilute form it is used in medicine. As a constituent of vinegar it is employed as a condiment, and in the preservation of foods.
- **Acetates** — Are salts produced by the partial or entire displacement of the hydrogen of acetic acid by a base. Sodic and lead acetates are of value in medicine. Iron, copper, and lead acetates are employed in the arts.

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QUESTIONS ON CHAPTER XXII.

1. What is the body which is present in vinegar which makes it sour? How would you obtain this substance from vinegar?

2. What are the chief properties of acetic acid? How would you demonstrate the same?

3. Describe the method of making vinegar by the quick process.

4. How are alcoholic liquids changed into vinegar?

5. What do you understand by the term acetous fermentation?

6 How would you prepare solid acetic acid?

7. Why is acetic acid classed as an organic body? State the components present, giving the proportion of each in 100 parts by weight.

8. What are acetates?

9. Give two examples of acetates, and say how you would prepare them.

10. What are the uses of acetic acid and acetates?

Preparation for Twenty-third Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of argol, tartaric acid, crust of wine, cream of tartar, tartar emetic, and calcium tartrate. Iron sand baths, tripod stand, litmus paper, lime-water, flask fitted with delivery tube, tumbler, test-tubes, sherbet, seidlitz powder, and other powders obtainable, which are sold for making effervescing draughts.

EXPERIMENTS TO BE PERFORMED.

"Show solubility of the solid acid in water, and that the solution has acid properties and is without odour.

Demonstrate the presence of carbon in the acid by ignition.

Show that argol also contains carbon.

Show how tartaric acid may be prepared from argol.

Prove that the action of a solution of tartaric acid in water results in the liberation of carbon dioxide from soda carbonate.

Prepare sherbet.

Mix the components of a seidlitz powder and show effervescence in water."—*Science Department's Syllabus.*

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CHAPTER XXIII.

TARTARIC ACID.

Occurs in many Fruits, especially in Grapes—Is obtained from Argol, an impure Potassium Salt of Tartaric Acid, deposited when Grape Juice ferments—Tartaric Acid is a crystalline Solid, and dissolves easily in Water—Has no smell—Is composed of Carbon, Hydrogen, and Oxygen —i.e., *the same Elements as are in Acetic Acid, but in different proportions—viz.,* 32·0 *parts of Carbon,* 4·0 *parts of Hydrogen, and* 64·0 *parts of Oxygen—Its action on Sodium Carbonate—Effervescing Draughts; Seidlitz Powders—Tartrates.*

TARTARIC ACID is the sour body which exists with **malic acid** in the juice of the grape, the pine-apple, the tamarind, and some other fruits. Mixed with other acid bodies it is found in the mulberry, and in the leaves of sorrel and other sour-leaved plants. Generally malic acid predominates in unripe fruit; and tartaric acid is more commonly met with in ripe fruit. For commercial purposes it is derived from grapes at different stages in the preparation of wine. A salt known as acid tartrate of potash exists in the juice of the grape. This salt is but slightly soluble in water, and even still less soluble in dilute alcohol; it is therefore deposited as a crystalline crust upon the surfaces of vessels in which grape juice is fermented. The deposition goes on even after the wine is put into bottle; that produced in the former case is known

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as **argol,** and that which is yielded in the latter is termed **crust.** Both deposits are also known as **tartar.**

Argol is *impure* acid tartrate of potassium, for it contains in addition to this salt varying quantities of tartrate of calcium and colouring matters. When purified by crystallization it is known as **cream of tartar.** From this body tartaric acid is prepared by first mixing with boiling water and then adding powdered chalk ; by this means **tartrate of lime** is produced. This substance is washed and then boiled with dilute sulphuric acid, which decomposes the salt of lime, fixing the base and liberating tartaric acid in solution. The liquor is next filtered to remove all the sulphate of lime ; the clear filtrate is concentrated by evaporation, and the tartaric acid obtained by crystallization. The crystals are still further purified by redissolving in water and recrystallizing.

Experiment 202.—Boil about 15 grams of argol with 75 cc. of water ; add powdered calcium carbonate as long as any effervescence is produced. Then add about 10 cc. of a strong solution of calcium chloride.

Observe that the liquid becomes turbid and a white precipitate is formed. This should be well washed with distilled water by decantation. Then prepare a mixture of, water 50 cc. and sulphuric acid 6 cc., add this to the calcium tartrate ; boil and filter. Evaporate the clear filtrate, then allow to cool, and observe the crystals that are formed.

We learn from this that tartaric acid may be prepared from argol by first converting it into **calcium tartrate,** and then decomposing this by means of dilute sulphuric acid; by which the calcium is fixed as the insoluble sulphate.

Tartaric acid is an inodorous, colourless, crystalline solid which will dissolve easily in hot water, but only slowly in cold water. The solid body has a sharp, pleasant, sour taste, which properties are also possessed by its

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solution. Tartaric acid is composed of the same chemical elements as acetic acid,—viz., **carbon, hydrogen,** and **oxygen**—but they are united in different proportions; 100 parts by weight of tartaric acid contain 32·0 parts of carbon, 4·0 parts of hydrogen, and 64·0 parts of oxygen.

Experiment 203.—Place a small crystal of tartaric acid on an iron plate, supported on a tripod stand, and heat it.

Observe that the mass chars and gives off the characteristic fumes and odour of burnt sugar.

We learn from this that tartaric acid contains carbon.

Experiment 204.—Introduce a small quantity of argol or crust of wine into a test-tube and strongly heat the same.

Observe that steam is evolved, and that a residue of carbon remains.

We learn from this that argol is an organic compound, for it contains the element carbon.

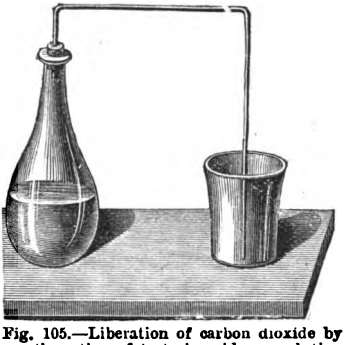

Fig. 105.—Liberation of carbon dioxide by the action of tartaric acid on a solution of sodic carbonate.

Experiment 205.—Expose a crystal of tartaric acid to the air for some hours or days.

Observe that there is no apparent difference after exposure, but that the crystal retains its clean bright surface.

We learn from this that the crystals of tartaric acid when exposed to air do not absorb water and *deliquesce* like calcium chloride, nor do they give up water and *effloresce* like sodic carbonate.

Experiment 206.—Place a crystal of tartaric acid in water in a test-tube; shake well, allow to stand for a time, and then heat. Taste the solution. Test the liquid with blue litmus.

Observe that the crystal dissolves slowly in cold water, and that it disappears at once in the hot water. The solution is sour or acid in flavour and will redden blue litmus.

We learn from this that tartaric acid is more soluble

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in hot than in cold water, that it has an acid reaction, and excites the sensation of a pleasant sour taste.

Experiment 207.—Fit up a piece of apparatus, as shown in Fig. 106. Place some sodic carbonate in the flask and pour on to it some of the solution of tartaric acid prepared in the last experiment; readjust the cork and tube so that the escaping gas may pass through lime-water.

Observe that on the addition of the tartaric acid to the carbonate of soda effervescence occurs, and the lime-water becomes milky.

We learn from this that tartaric acid, like acetic, is a stronger acid than carbonic, for it will displace the carbonic dioxide for sodic carbonate. The salt thus formed is **sodium tartrate**; other tartrates may be prepared in a similar manner by the action of the acid upon carbonates.

EFFERVESCING DRAUGHTS.—Most of the solid powders which are employed in the preparation of such drinks consist essentially of carbonate of soda and tartaric acid. They are sometimes rendered more palatable by the addition of sugar and other substances. Some such mixtures are artificially coloured, and reach the consumer under highly fanciful names.

Experiment 208.—Grind well together in a mortar a mixture consisting of one dram of solid bicarbonate of soda and three-quarters of a dram of tartaric acid. Add this to five or eight ounces of water.

Observe that the two bodies do not act chemically upon each other in the solid state, but that on the addition of water the substances both dissolve with effervescence. When the liberation of gas has quite ceased notice that the remaining liquid has a decided saline flavour.

We learn from this that an effervescent mixture may be prepared from sodic carbonate and tartaric acid. The palatability of this substance may be increased by the addition of ground loaf sugar and essence of lemon. A mixture thus prepared is termed **sherbet**, or **lemon kali**.

Seidlitz powders are usually put up in two small packets. One contains 45 grains of sodium bicarbonate

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mixed with 120 grains of Rochelle salt, and the other about 35 grains of powdered tartaric acid. For use the former powder is first dissolved in water, the acid is next added; this sets free carbon dioxide from the carbonate of soda, causing violent effervescence. The solution forms an agreeable and cooling drink which is mildly aperient in its action.

Tartrates are salts which are formed by the union of a base with tartaric acid; many are of use in medicine.

Tartar emetic is the best known of the salts; it is prepared by dissolving antimonious oxide in cream of tartar.

Cream of tartar is the acid tartrate of potash.

Rochelle salt is sodium potassium tartrate.

SUMMARY TO CHAPTER XXIII.

TARTARIC ACID

- **Preparation** — Occurs free in sour fruits, and in the form of salts. Made from *argol*, the crude and impure acid tartrate of potash produced during the fermentation of grape juice. Argol by recrystallization is purified and yields *cream of tartar*. This is first converted into tartrate of lime, and then decomposed by means of sulphuric acid.
- **Properties** — It is a transparent, inodorous, crystalline solid with a sour taste, and soluble in water. It reddens litmus and decomposes carbonates with effervescence. Tartaric acid with soda carbonate forms sodium tartrate, and liberates carbon dioxide. Since the effervescence lends the drink a pleasant palatability, and the sodium salt produced is mildly aperient, this mixture is employed in the preparation of cooling drinks.
- **Uses** — Tartaric acid and tartrates are used in medicine also as a substitute for citric acid and lemon juice in the preparation of cooling beverages. Tartaric acid is also employed in calico printing for the purpose of removing colours from some portion of the material.
- **Tartrates** — Salts produced by the partial or entire displacement of the hydrogen of tartaric acid by a base. The more important are *cream of tartar*, acid tartrate of potassium, and *tartar emetic*, tartrate of antimony.

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QUESTIONS ON CHAPTER XXIII.

1. How is tartaric acid found in nature?

2. Describe the preparation of tartaric acid from grape juice?

3. What is the composition of tartaric acid?

4. Describe fully how you would demonstrate the chief properties of tartaric acid.

5. How would you prove that carbon dioxide is produced by dissolving a seidlitz powder in water?

6. Describe the composition of sherbet.

7. Name any tartrates, and describe their composition.

8. What is cream of tartar?

9. What is argol? How is it prepared? What is made from it?

10. How would you prove that argol, cream of tartar, and tartaric acid all contain carbon?

Preparation for Twenty-fourth Lesson.

APPARATUS AND MATERIALS REQUIRED.

Specimens of palmitin candles, almond, olive, linseed, hemp, poppy, colza, castor, cod-liver, and neatsfoot oils, mutton suet, tallow, and butter. Evaporating dishes, watch-glasses, test-tubes, beaker, sodic hydrate, potassic hydrate, and alcohol.

EXPERIMENTS TO BE PERFORMED.

Demonstrate the chief differences between volatile and fixed oils.

Show how drying oils and non-drying oils differ in properties.

"Tie beef or mutton fat up in a muslin bag, and melt in boiling water to separate the fat from membraneous matter.

Show that fat is insoluble in water, that it floats on water, and melts at a temperature below boiling water.

Show that oil has very similar properties to melted fat.

Boil oil or fat with caustic potash.

Prepare a solution of potassium stearate, and precipitate stearic acid from it by the addition of hydrochloric acid.

Show the solubility of the acid in alcohol and ether, and the insolubility of the lime salt of stearic acid.

Obtain crystals of stearic acid by evaporating an alcoholic solution of the acid."—*Science Department's Syllabus.*

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CHAPTER XXIV.

FATS AND OILS.

Are neutral Bodies made up of an Acid and a Base, the Base in all cases is Glycerine; the Acid varies in different Oils and Fats—They are all insoluble in Water—Oils are Liquid; Fats are Solid—Many of the Oils are obtained from Vegetables, either from the Seed or Fruit—Most of the Fats are obtained from Animals—Melting of Tallow (Fat of the Ox, Sheep, etc., put in Boiling Water)—Its non solution in Water—Its lightness as compared with Water—If Fat be boiled with a Solution of Caustic Potash, the Liquid becomes slightly Milky, and nearly the whole of the Fat dissolves—Combination of the Potash with the Acid (Stearic) of the Tallow and formation of Potassium Stearate—Previously the Stearic Acid was combined with Glycerine—To the Solution of Potassium Stearate Hydrochloric Acid is added—The Potash is again separated from the Stearic Acid, and the Stearic Acid, as it cannot dissolve in Water, separates out—Stearic Acid dissolves in Alcohol and in Ether, and separates out in Crystals on Evaporation—Used in making Candles, and is better than Tallow because it melts at a higher Temperature—Tallow distilled with Steam of Temperature 600° F. (high pressure Steam) separates into Stearic Acid and Glycerine, and when cold these Bodies remain separate—All Oils and Fats are decomposed by Potash in the same way as Tallow.

OILS AND FATS are neutral bodies made up of an acid and a base; the base in all cases is **glycerine**, but the acid present varies with different oils and fats. *The fats and fixed or greasy oils* may be looked upon as **organic salts**, in all of which glycerine is present as a base; such bodies have therefore been called "**glycerides.**"

The most important "glycerides" are *stearin*, *palmitin*, *olein*, and *linolein*, which are compounds of glycerine with

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stearic, *palmitic*, *oleic*, and *linoleic acids* respectively. They are all insoluble in water; the oils are liquids, and the fats are solids at ordinary temperatures.

Common oils are usually composed of varying quantities of olein and palmitin; thus **almond oil** consists of olein with a small quantity of palmitin; **olive oil** of palmitin and olein; **linseed oil**, **hemp seed oil**, and **poppy oil**, all contain linolein. This body is present also in the fatty exudations of animals, and wool is charged with it. The linolein of linseed oil absorbs oxygen on exposure to air, and solidifies, and thus causes the oil to "dry."

The ordinary fats, such as **tallow**, **lard**, and **suet**, are mixtures of stearin, olein, and palmitin, in various proportions; the softer with low melting points contain much olein, and the harder and less fusible contain most stearin. Stearin is the solid present in lard and mutton suet; palmitin is the semi-solid body present in the human fat, and palm oil. Stearin and palmitin melt at 160° F. and 140° F. respectively. Linolein and olein are liquids.

OILS are classified as **volatile** or **essential**, and **fixed** or **fatty**. The former class includes oil of cloves, oil of lavender, etc. Olive, colza, linseed, palm-oil, and tallow belong to the latter class. Essential oils are distinguished from fixed oils by the following properties :—

1. *Volatile oils evaporate readily* and do not leave a stain on linen or paper, whilst *fixed oils* make a *translucent grease spot*, which is not removed by warming.

2. *Fixed oils soil the skin* and produce a greasy or **oleaginous** sensation when rubbed with the fingers, whilst *volatile oils do not produce the greasy sensation*.

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3. The *volatile oils* are *more inflammable than fixed oils.*

4. The *volatile oils* are all *more fluid than fixed oils* under the same conditions.

5. The *fatty oils* have usually *only a very slight odour and faint taste*; on the other hand the *essential oils are aromatic* and possess very *decided flavours.*

Many of the oils are obtained from the vegetable kingdom, either from the fruits or seeds of plants; most of the fats on the other hand are of animal origin. The **mineral oils** like **paraffin,** and the solids, **vaseline** and **paraffin wax,** belong to a totally different class of bodies.

Experiment 209.—Mix with some water in a test-tube some colza oil, and with a second portion of water in another tube mix some oil of turpentine. Shake both well.

Observe that the oils separate out on standing.

We learn from this that oils are insoluble in water.

Experiment 210.—Pour into two dishes colza oil and turpentine.

(*a*) Dip the fingers first into the colza oil, wash off the oil, and then test the turpentine in a similar manner.

(*b*) Place a drop of each liquid on a piece of blotting paper by means of a glass rod, and warm the paper in each case.

(*c*) Pour a few drops of each liquid into two test-tubes, and heat the tubes separately in a Bunsen flame.

(*d*) Introduce a very small quantity of each of the oils into two evaporating dishes. Plunge a burning taper into each.

Observe that colza oil has only a faint odour, and is a thickish liquid; whilst turpentine has a decided odour, and is a clear, colourless, limpid liquid.

(*a*) The colza oil is greasy to the touch, whilst turpentine is not oleaginous.

(*b*) Colza oil leaves a greasy stain on the paper because it is not volatile, whilst the turpentine completely dries up and does not soil the paper.

(*c*) The turpentine is evaporated by heat, but the colza oil remains, and after a time chars.

(*d*) The colza oil does not readily burn, but extinguishes the taper; but the turpentine is inflammable.

We learn from these experiments the chief differences between colza oil, which belongs to the class of fatty or fixed oils, and turpentine, which ranks as an essential oil.

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Since these are types of classes, we thus learn the important distinguishing properties of each class.

Essential oils are obtained by distilling the leaves and petals of the flowers of plants with water. When such oils are dissolved in alcohol they furnish **essences.**

Fixed or fatty oils, according to their behaviour when exposed to air, are divided into the two classes—

1. **The drying oils,** which include the following :—
   (*a*) *Linseed oil,* used in making paint, varnish, and enamel.
   (*b*) *Hemp oil,* employed in the preparation of varnish.
   (*c*) *Poppy oil,* used in the manufacture of soap and varnish.
2. **The non-drying oils,** which include the following :—
   (*a*) *Colza or rape oil,* employed as a lubricant and for illuminating purposes.
   (*b*) *Olive oil,* used for lamps and as a food.
   (*c*) *Cotton-seed oil,* used as a lubricant and in making soap.
   (*d*) *Ground-nut oil,* used in making soap.
   (*e*) *Almond oil,* used in medicine.
   (*f*) *Castor oil* „ „
   (*g*) *Codliver oil* „ „
   (*h*) *Neatsfoot oil,* liquid at low temperatures, and therefore used for lubricating clocks and delicate machinery.

Experiment 211.—Expose in two shallow dishes specimens of *olive oil* and *linseed oil.* Allow them to remain unprotected from the air for several days.

Observe that the olive oil undergoes no apparent change, but that the surface of the linseed oil becomes covered with a thin skin.

We learn from this that olive oil does not dry when exposed to air, but under the same circumstances linseed oil behaves as a drying oil.

The drying oils absorb oxygen from the air, and form resins which cause them to solidify; but the non-drying oils, after a time, develop acids and become rancid.

FATS are those **oleaginous bodies** which remain solid at the ordinary temperatures of the air. Butter, lard, and suet or tallow, are the common fats. If mutton suet is

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placed in hot water much of it melts and floats to the surface as a clear, colourless, oily mass; this solidifies on cooling, but remains on the surface of the water.

Experiment 212.—Tie a few fragments of mutton fat in a loose muslin or old linen bag and immerse it in boiling water.

Observe that the fat slowly melts out and floats to the surface.

We learn from this that mutton fat is lighter than water, and that it melts at a lower temperature than that of boiling water. The stringy substance which remains in the bag consists of the fibre and membrane which bound the fat together into suet.

Action of alkalies on fats.—If some of the fat obtained in the last experiment be boiled with a solution of caustic potash the liquid becomes slightly milky, and nearly the whole of the fat dissolves. This result is due to the fact that the strong base, caustic potash, decomposes the fat, driving out glycerine, and combining itself with the fatty acid. In the case of tallow, which consists mainly of **stearin**, the potash combines with the **stearic acid**, forming **stearate of potash**, and liberates glycerine. This stearate of potash is really **soft soap**; from this the potash may be removed by the acid of hydrochloric acid. If some hydrochloric acid be added to a solution of stearate of potash, chloride of potassium is formed, and the **stearic acid** separates, and, as it is lighter than water and insoluble in that liquid, it floats to the surface.

Stearic acid is soluble in ether and in alcohol, and when an alcoholic solution of stearic acid is allowed slowly to evaporate it yields *crystals of stearic acid*; this is a white solid body which is harder than tallow, and melts at a higher temperature than that body; it is, therefore,

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employed in making hard candles. Stearic acid may also be obtained from tallow by the action of steam under pressure; for at a temperature of about 350° C. the stearin breaks up into stearic acid and glycerine.

All oils and fats are decomposed by potash in the same way as tallow, and the following equation will make the reactions clear to the student:—

$$\left.\begin{matrix}\text{Stearate}\\ \text{Oleate}\\ \text{Palmitate}\end{matrix}\right\}\text{of glycerine} + \left\{\begin{matrix}\text{caustic soda}\\ \text{or}\\ \text{caustic potash}\end{matrix}\right\} = \left\{\begin{matrix}\text{Stearate}\\ \text{Oleate}\\ \text{Palmitate}\end{matrix}\right\}\text{of soda or potash} + \text{glycerine.}$$

Experiment 213.—Boil some stearin obtained in experiment 212 in an evaporating dish with a fairly strong solution of caustic potash. From time to time add, drop by drop, a strong solution of caustic potash.

Observe that the liquid becomes milky and the fat mainly disappears, and when the liquid cools a soft mass of stearate of potash remains.

We learn from this that fat will dissolve to a considerable extent in caustic potash to form a somewhat milky liquid which on cooling yields **soft soap.**

Experiment 214.—Dissolve a portion of the potassium stearate, obtained in the last experiment, in distilled water by gently heating; when a clear liquid is obtained, carefully add hydrochloric acid.

Observe that a white solid waxy-looking mass of stearic acid floats to the surface of the cold liquid.

We learn from this that stearic acid may be obtained from stearin by the action of hydrochloric acid.

Experiment 215.—Place a small fragment of stearic acid obtained in the last experiment in a test-tube, and add alcohol; shake well.

Observe that the stearic acid disappears, and, on standing, separates out in a crystalline form.

We learn from this that stearic acid is soluble in alcohol. This experiment may be repeated with ether, and we should then prove that this liquid will also dissolve stearic acid.

Experiment 216.—Dissolve a small portion of the soft soap obtained in experiment 213 in distilled water; add a solution of lime-water.

Observe that a whitish scum separates, and forms flakes in the water.

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We learn from this that lime converts the soluble soap, stearate of potash, into an **insoluble soap, stearate of calcium.**

Common hard soap is stearate of soda, and this forms an insoluble soap; when added to hard waters, the scum of matter produced is stearate of lime, or magnesia.

Stearate of soda or potash. (*Soluble soap.*) + Salts of lime or magnesia. = Stearate of lime or magnesia. (*Insoluble soap.*) + Salts of soda or potash.

SUMMARY TO CHAPTER XXIV.

FATS -
- Salts consisting of a fatty acid combined with glycerine; hence they are spoken of as **glycerides.** The most important are:—
- *Stearin*, which solidifies at 72° C.
- *Palmitin* „ „ 66° C.
- *Olein* „ does not solidify at 0° C.
- They are all insoluble in water and neutral. They are decomposed by caustic soda or potash, or by great heat into fatty acids and glycerine.

OILS -
- Are those fats which remain liquid at ordinary temperatures.
- *Fixed or fatty*, those which are not volatile.
- *Essential*, those which are easily volatilized.
- *Drying oils*, those which form a resin and harden in the air.
- *Non-drying oils*, of value for lubricating and illuminating purposes; they do not harden or dry when exposed to air.

FATTY ACIDS
- These bodies are liberated from fats by the action of heat or caustic alkalies. Thus potash acting upon—
- (*a*) Stearin liberates *stearic acid.*
- (*b*) Palmitin „ *palmitic acid.*
- (*c*) Olein „ *oleic acid.*
- With these the alkali combines to form a **soap.** We therefore see that potash combines with—
- Stearic acid to form stearate of potash.
- Palmitic „ „ palmitate „
- Oleic „ „ oleate „
- From these salts (soaps) the fatty acid may be liberated by the action of hydrochloric acid.

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QUESTIONS ON CHAPTER XXIV.

1. To what class of substances do fats belong?

2. What is the difference between an oil and a fat?

3. Give the chemical names of the common fats.

4. What fatty substance is present in mutton suet?

5. What are the chief physical properties of fats?

6. How may oils be classified? How would you distinguish an essential from a fixed oil?

7. For what purposes are oils employed? Give examples.

8. Why do some oils dry? Give examples of drying and non-drying oils.

9. How would you obtain stearin and stearic acid from mutton fat?

10. What are soaps? Describe fully the chemical changes which take place when caustic alkalies act upon fats.

Preparation for Twenty-fifth Lesson.

APPARATUS AND REAGENTS REQUIRED.

Glycerine, test-tubes, olive oil, litharge, evaporating dish and water bath, sulphuretted hydrogen apparatus, and acetic acid.

EXPERIMENTS TO BE PERFORMED.

Demonstrate the solubility of glycerine in water and its sweet taste.

Prepare glycerine from olive oil.

Prepare a glyceride by the action of acetic acid upon glycerine.

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CHAPTER XXV.

GLYCERINE.

A thick colourless Liquid with a sweet taste—Dissolves readily in Water—When quite pure becomes solid at a low Temperature—If heated alone it is destroyed, but if heated with Water in a Retort it distils over with the Steam—Heated with Acids it combines with them, and Bodies similar to Fats are formed.

GLYCERINE is a sweet, thick, colourless liquid, which dissolves readily in water. It is somewhat heavier than water, having a specific gravity of 1·27. When pure it solidifies at a very low temperature ; it has therefore been employed in the construction of barometers, for it is much lighter than mercury, and thus shows readily a slight variation in atmospheric pressure. Glycerine is largely prepared as a by-product in the manufacture of soap by the decomposition of fat. Through the agency of superheated steam, fat is broken up with liberation of glycerine. In the preparation of soap the distillation is carried on at high temperatures under pressure, and the liquid which condenses in the receiver separates into two layers, one consisting of the fatty acid and the other of glycerine. When heated with acids glycerine unites with them, forming salts. Thus if glycerine be mixed with pure acetic acid and heated, combination takes place, and a series of **glycerides** are

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produced which vary in composition with the proportion of the ingredients used and the temperature employed. These bodies are similar in chemical constitution to fats, in fact, they belong to the same class of substances, viz., glycerides.

When heated alone, glycerine boils at 290° C., and undergoes decomposition ; but if heated with water in a retort it distils over with the steam.

Glycerine is employed for medicinal purposes, and in the preparation of cosmetics and pomades ; it is also used for the preservation of anatomical specimens and microscopical sections. Treated with nitric and sulphuric acids it forms the highly explosive and poisonous liquid known as *nitro-glycerine*. When sawdust, earth, etc., are saturated with nitro-glycerine the body termed *dynamite* is produced.

Experiment 217.—Pour some glycerine into a test-tube, and then add water drop by drop, and shake the mixture after each addition.

Observe that the glycerine will mix with water in all proportions.

We learn from this that glycerine is readily soluble. So great is the affinity of glycerine for water that when exposed to air it absorbs moisture and increases in weight.

Experiment 218.—Place in an evaporating dish about 100 grams of *olive oil*, add about three times the volume of water, then heat the mixture over a water bath, and stir in gradually about 50 grams of finely-powdered *litharge*, oxide of lead. Continue to heat the mixture with frequent stirring for several hours.

Observe that after a time a paste-like mass is obtained of **lead soap.** The clear liquid should be poured off at this stage, and diluted with an equal volume of water ; sulphuretted hydrogen should be passed through the liquid to precipitate any lead salt. The liquid should next be filtered to remove any insoluble sulphide of lead, and then the clear filtrate should be carefully evaporated on a water bath. A sweet liquid is left, as the solution is concentrated, which is glycerine.

We learn from this that glycerine may be prepared from the olein of olive oil by first fixing the oleic acid by

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means of lead, as **lead oleate, or lead soap**; thus the glycerine with which the acid was combined is liberated and remains in solution in the water. Any excess of uncombined lead oxide is removed by sulphuretted hydrogen, and the aqueous solution of glycerine is concentrated by the evaporation of the water.

Experiment 219.—Prepare a mixture of pure glacial acetic acid and glycerine and heat it in an evaporating dish to about 100° C. Pour the liquid when cold into water.

Observe that the product is unlike both glycerine and acetic acid.

We learn that the organic acid, acetic, will combine with glycerine; the latter body, therefore, behaves as a base. The substance which we have produced in this experiment is a **glyceride**, and it is built up on a similar plan to that upon which all the fats are constituted. By varying the proportion in which the substances are mixed, and the temperature to which the mixture is raised, we may obtain different products all allied in constitution to fats and oils.

SUMMARY TO CHAPTER XXV.

GLYCERINE

- **Properties** - A thick, colourless liquid with a sweet taste; dissolves readily in water in all proportions. Combines with acids to form glycerides. Fats are glycerides in which fatty acids are united with glycerine.
- **Preparation** - This body is obtained during the manufacture of soap, in the preparation of which fat is decomposed by superheated steam.
- **Uses** - - - It is employed for medicinal purposes, in the preparation of cosmetics, and in the preservation of botanical, anatomical, and histological specimens.

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QUESTIONS ON CHAPTER XXV.

1. How is glycerine met with in nature?

2. Describe the chief properties of glycerine.

3. How is glycerine prepared?

4. Explain fully what experiments you would perform in order to demonstrate the chief properties of glycerine.

5. What are glycerides? Name three common glycerides. Explain how you would obtain glycerine from the body.

6. What is lead soap? How may this body be prepared?

7. Explain how you would prepare any glyceride.

8. How would you prove that glycerine is soluble in water?

9. What are some of the common uses of glycerine?

10. How would you prove that glycerine contains carbon?

Preparation for Twenty-sixth Lesson.

APPARATUS AND REAGENTS REQUIRED.

Specimens of hard soap, soft soap, marine soap, Marseilles soap, various kinds of toilet soaps, beef fat or mutton suet, caustic soda, caustic potash, common salt, alcohol, distilled water, hard water, test-tubes, filter, filter papers, lime-water.

EXPERIMENTS TO BE PERFORMED.

Shake distilled water up in bottle with soap and solutions of salts, and show action of acids upon the solution of soap.

Add soap solution to distilled water, also to common water, and explain the difference of action.

Show the presence of stearic acid in soap by adding hydrochloric acid to a solution of soap.

Show how hard water may be softened by the use of lime.

Illustrate method of preparing soap from fat and caustic soda.

Demonstrate the method of determining the hardness of water by the soap test.

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CHAPTER XXVI.

SOAP.

By boiling Fat with Caustic Soda Sodium Stearate is formed—On adding Salt to the Liquid the Sodium Stearate, which is Soap, separates out and solidifies on the Surface of the Liquid—Soft Soap is Potassium Stearate—Action of Soap in washing—Action of Soap on Hard and on Soft Water.

SOAP is an organic chemical compound which belongs to the class of bodies which we have learned to call salts. There are several chief varieties of soap, but all agree in that they are produced by the action of an alkali upon a fat. In the two last lessons we have seen that fats are split up into fatty acids and glycerine at high temperatures or by the action of caustic alkalies. When caustic soda or caustic potash acts upon oils or fats the soda or potash salts which are formed of the fatty or oily acids are **soaps.** This process is called "**saponification,**" or soapmaking, and a fat or oil is said to be **saponified** when it is acted upon by a caustic alkali. In order to separate the soap from the water in which it is formed, common salt is added to the liquid, and, as soap is insoluble in brine, it separates out as a curdy solid, leaving glycerine in solution.

Common hard soap is *stearate of soda.*

Experiment 220.—Cut up some pieces of beef or mutton *fat* and throw them into a small quantity of boiling water contained in a beaker. The volume of water should not be more than four times that of the fat.

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When the fat is melted add about two volumes of a strong solution of *caustic soda* and heat over a water-bath with frequent stirring. When saponification is complete add about half a volume of a strong solution of common salt.

Observe that the fat melts out in the boiling water, and that it is saponified by the action of caustic soda. Notice that on the addition of the common salt a curd separates out and floats upon the mixture. When the liquid is cold this curd of imperfect soap should be removed and washed carefully with water; it should next be placed in a little warm water and caustic soda added, drop by drop, with constant stirring, until the mass is completely dissolved. When saponification is perfect the liquid should be allowed to cool; a more or less solid soap will then separate out.

We learn from this that **sodium stearate** may be produced by boiling fat with caustic soda. On the addition of salt the impure soap separates; this may be further saponified and converted into a more perfect soap by dissolving in water with the addition of more caustic soda.

Soft soap is mainly *stearate of potash*. It is prepared by boiling crude potash with fish or other oil and tallow.

Marine soap is made by saponifying cocoanut oil. It derives its name from the fact that it may be used for washing in sea-water, because it is much more soluble in salt water than ordinary soap.

Marseilles soap is the marbled variety of soap which is made, chiefly in the south of Europe, from olive oil.

Most common soaps contain free alkali; those sold for toilet purposes contain but little, for they are usually prepared with great care, so keen is the competition between rival manufacturers. Colouring matters are often added to please the eye or to mask an inferior soap, and perfumes are sometimes employed to hide an unpleasant odour.

The cleansing power of soap depends upon the fact that it possesses the property of dissolving fatty substances. In washing the skin we know that water alone is

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not effectual. Fatty matter is continually secreted from the blood by little glands in the skin, and this fat softens the skin, exudes from the glands, and accumulates upon the surface. This oily matter is insoluble in water, but is easily removed by the action of soap.

Good white soap is soluble in dilute alcohol and the solution is not precipitated by distilled water, but if added to hard water or to a solution containing a calcium salt double decomposition takes place and an **insoluble lime soap** is produced. The soluble stearate of sodium of ordinary washing soap is by this means converted into insoluble stearate of calcium. London water, and all other waters which contain lime and magnesia salts in solution, are for this reason unsatisfactory for washing. The insoluble lime or magnesia compound, produced by the addition of common soap to such waters, forms a curd which tends to adhere to the surfaces of materials which are washed in them. It is only after the limy matters of hard waters have been removed that the soap will dissolve, and that its cleansing properties become apparent. But the softening of hard water, by the use of an excess of soap, is a wasteful and troublesome method. For domestic and manufacturing purposes water should therefore be softened by boiling, or by the addition of lime, soda, or potash.

In the process for the **determination of the hardness of water,** which we owe to the late Dr. Clark of Aberdeen, and which is known as **Clark's soap test,** we take advantage of the following facts :—

(*a*) That soap will dissolve with ease in alcohol.
(*b*) That it is converted into an insoluble compound by lime or magnesia.

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(*c*) That it will not form a lather with hard water until the hardening salts have first been converted by it into insoluble soaps.

For practical purposes an alcoholic solution of soap is prepared of known strength, and the amount of this required to produce a lather with a given quantity of the water under examination affords us the data from which the degree of hardness may be calculated.

Experiment 221.—Add some shavings of ordinary soap to distilled or rain-water in a flask, and well shake the mixture.

Observe that the soap at once dissolves and forms a lather.

We learn from this that soap will dissolve at once in rain or distilled water.

Experiment 222.—Place about the same quantity of soap as was used in the last experiment in about the same volume of alcohol, and shake.

Observe that the soap dissolves in the spirit at once.

We learn from this that soap is also soluble in alcohol.

Experiment 223.—Place in two bottles equal volumes of distilled water and ordinary hard drinking water; to each add a few drops of the alcoholic solution of soap prepared in experiment 222, and shake each well.

Observe that a lather is formed at once with distilled water, but that a much larger quantity of soap solution is required to produce a lather with hard water. By the addition of more soap solution the hard water may be made, in time, to yield a good permanent lather.

We learn from this that much more soap is required in order to produce a lather with hard water than is necessary for the same quantity of soft water.

Experiment 224.—Add to some lime-water a few drops of the alcoholic soap solution prepared in experiment 222, and shake well.

Observe that the soap forms a white curd, which floats in the liquid as flakes, and rises to the surface as a scum on standing.

We learn from this that lime-water curdles soap forming the insoluble lime-soap, **calcium stearate.**

Experiment 225.—Add to some ordinary hard London water about an equal volume of clear lime-water, allow to stand some time, then filter.

Observe that a fine white substance is thrown out of solution.

We learn from this that the **soluble bicarbonate of lime** of London water may be converted into the **insoluble carbonate of lime** by the action of lime-water.

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Experiment 226.—To a portion of this water, equal in bulk to that employed in experiment 223, add about the same quantity of soap solution.

Observe that a lather forms at once, showing that the water is soft.

We learn from this that some hard waters may be softened by the action of lime-water. Such a method of **softening hard waters** is employed in many cases on a large scale, and these softened waters are better and always more economical for domestic purposes than hard water.

Experiment 227.—Shake up some of the solution of soap prepared with distilled water in experiment 221 with hydrochloric acid.

Observe that a solid substance forms flakes which float on the liquid.

We learn from this that hydrochloric acid will decompose soap with liberation of stearic acid. In this experiment the chlorine of the acid combines with the base present in the soap and forms sodium chloride.

SUMMARY TO CHAPTER XXVI.

SOAP

- **Composition** — Chemical compound produced by the action of an alkali upon a fat. Common soap is a compound of stearic acid and soda.
- **Preparation** — Common soaps prepared by decomposing fats at high temperatures into their fatty acids and glycerine and then causing the fatty acid to combine with soda or potash.
- **Properties** —
  - Soluble soaps will dissolve in water and alcohol; insoluble soaps cannot be dissolved by these liquids.
  - Soda and potash soaps will dissolve in water, and since they contain a small amount of free alkali they assist in the cleansing of oily surfaces.
- **Kinds** —
  - *Common hard soap:* Stearate of soda; soluble in water and alcohol.
  - *Soft soap:* Stearate of potash; very soluble, and alkaline in reaction.
  - *Marine soap:* Prepared from cocoanut oil; employed for washing with salt water.
  - *Marseilles soap:* made from olive oil and soda.
  - *Toilet soaps:* Sodium stearate, to which various colouring matter and perfumes are often added.
  - *Lime and magnesia soaps* are produced by the action of hard water upon ordinary soap.

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QUESTIONS ON CHAPTER XXVI.

1. What is the chemical composition of common hard soap?
2. To what class of bodies does soap belong? Give reasons for your answer.
3. What chemical elements are present in soap?
4. Give an outline of the method of preparing soap.
5. How does soft soap differ from common yellow soap?
6. What is marine soap, and how is it prepared?
7. What are the chief properties of soap, and how would you demonstrate the same?
8. Why does soap curdle in some kinds of water?
9. Describe Clark's soap test for hardness of water.
10. How would you prove that soap contains stearic acid?

Preparation for Twenty-seventh Lesson.

APPARATUS AND REAGENTS REQUIRED.

Specimens of white and brown sugar, barley sugar, sugar candy, treacle, golden syrup, molasses, honey, and glucose. Three beakers, lime, cloth filtering bag suspended from a stand as shown in Fig. 106, granular bone charcoal, litmus papers, sand bath, evaporating dish, stone slab, tripod stand, alcohol.

EXPERIMENTS TO BE PERFORMED.

Illustrate sugar refining by passing solution of raw brown sugar through woollen bag. Show decolourization of sugar by action of bone charcoal.

Show preparation of barley sugar and caramel.

Illustrate the colouring properties of caramel.

Show great solubility of cane sugar in water; also that its solution is neutral.

Heat some cane sugar and point out the peculiar odour it gives out, and that on further continuing the heat it leaves a residue of carbon.

Wash honey with spirit, and show the residue is sugar, but that it is not so sweet as ordinary sugar and not so soluble.

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CHAPTER XXVII.

SUGAR.

Exists in many Plants; is obtained from the Sugar Cane; also from Beet Root—The Juice of these Plants yields the Sugar; when pure it is White, Crystalline, Sweet, and very Soluble in Water—Sugar Candy—If heated with very little Water to 365° F., on cooling it is no longer Crystalline and is "Barley Sugar"—Does not combine with Acids, but even a very little Acid boiled for a long time with a Solution of Sugar changes to another kind of Sugar—Composition of Cane Sugar—The several different kinds of Sugar, e.g., the solid part of Honey is a Sugar which differs from the Sugar in the Sugar Cane—The same found in all sweet Fruits, and is called Grape Sugar—Grape Sugar not so sweet nor so soluble as Sugar Cane.

SUGAR occurs naturally in many plants. It is an important constituent of fruits and as such serves as a food substance. Sugar is so largely employed in the preservation and sweetening of foods, and in the preparation of beverages, that its manufacture from vegetable substances constitutes an important industry.

For chemical reasons sugars are divided into two classes, viz., the **sucroses** and the **glucoses**. The best known members of the former class are *cane sugar* and *milk sugar*, and varieties of glucose are present in *honey* and *fruits*.

Cane sugar is obtained chiefly from the juice of the sugar-maple, beet, sugar-palm, and sugar-cane.

To obtain sugar from the sugar-cane, the plants are cut down, removed from the field, and the canes are then broken

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into pieces and crushed between large iron rollers. The juices which run out are caught in vessels placed beneath the rollers. This liquid is heated in copper vessels until it becomes thick, when the required consistency has been attained it is poured out into open pans and allowed to cool. From this evaporated sap **brown sugar** crystallizes out, leaving behind a liquid known as **molasses**. When the mass is quite cold it is placed in large casks containing a number of small holes in the bottom, through which any portion of the juice which will not solidify runs out. The material which remains in the cask is known as **raw** or **moist sugar**, and the liquid which runs away is called **treacle**. The brown sugar thus prepared is still further purified before it reaches the market as **loaf sugar**. It is dissolved in water; a small quantity of albumin, usually in the form of blood, is then added; when this mixture is heated the albumin coagulates, and settling in the liquid carries down many of the impurities. For this stage of purification, lime is also sometimes employed. The saccharine liquid is clarified by filtering through bags made of woollen cloth or horsehair, and is next passed through animal charcoal, by which means the remaining colouring matter is removed. The colourless liquid is placed in large copper vacuum pans and concentrated at a low temperature. The pans are air-tight vessels provided with strong pumps, by which means the air and moisture are drawn out, and the liquid is made to boil at about 74° C. When sufficiently concentrated, the sugar is poured into moulds and allowed to cool gradually; the greater portion crystallizes out as **loaf sugar**, but a thin liquid remains which is known as **golden syrup**.

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The sugar obtained by the above method is a white, crystalline, sweet solid, which is very soluble in water. When a strong solution of brown sugar is allowed slowly to evaporate, it yields **brown sugar candy**, and if some white sugar be dissolved in water and treated in the same manner, **white sugar candy** is produced. If sugar is boiled with a small quantity of water until it assumes a slightly yellowish tinge, which change is produced at about 190° C., the resulting body when cold is no longer crystalline but glassy, and in this state it is known as **barley sugar.** On raising the temperature of the solution to a higher point, the liquid darkens and the sugar passes into **caramel**, a substance which is largely employed in falsifying wines and spirits, and for colouring soups and gravy.

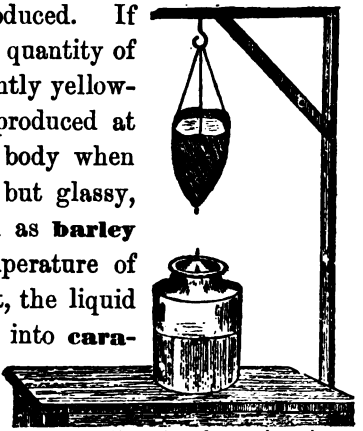

Fig. 106.—Filtration of sugar through cloth bag.

Milk sugar or **Lactose** is present in the milk of all mammals, and is obtained by evaporating whey at a low temperature; it then solidifies in the crystalline form.

Glucose or **invert sugar** is a substance which is found in ripe fruits, and in honey; it is less sweet than cane sugar, and not so soluble in water. From the fact that it is found in ripe grapes it is known as **grape sugar.** Ordinary cane sugar is easily converted into grape sugar by boiling it with dilute sulphuric acid. Glucose is also easily made from starch, or from the cellulose in cotton rags, sawdust, etc., by boiling with sulphuric acid. This variety of sugar is also produced from starch, during the germination of most seeds, by the action of a ferment which is developed during this period of its growth. In the bodies of animals starch is also converted into glucose by the action of special ferments.

Levulose is present in honey. It is a colourless, crystallizable syrup, sweeter and more soluble in water and alcohol than grape sugar.

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All the varieties of sugar consist of carbon, combined with hydrogen and oxygen, the latter bodies standing in the same proportion to each other as they do in water. For this reason sugars are classed with starch and some other substances as **carbo-hydrates.** Sucroses differ from glucoses in that they contain by weight more carbon in proportion to the hydrogen and oxygen present.

Experiment 228.—Dissolve some common raw or brown sugar in warm water, mix it with a little powdered lime, and pass it through a filter made of thick cloth in the form of a bag as shown in Fig. 106.

Observe that a clear but slightly discoloured syrup runs through.

We learn from this that the first stage in **sugar refining** may be brought about by the action of lime.

Experiment 229.—Pass some of the discoloured syrup obtained in the last experiment through a filter of loosely packed granular bone charcoal.

Observe that a clear colourless liquid runs through. This should be carefully concentrated over a water-bath, and then placed aside to cool.

We learn from this that charcoal will decolourize the syrup, and from this sugar may be obtained by evaporation and crystallization. The last two experiments roughly illustrate the method of obtaining loaf sugar from brown sugar.

Experiment 230.—Dissolve as much sugar as possible in a small beaker of distilled water. Test the liquid with red and blue litmus papers.

Observe the extreme solubility of sugar, the sweetness of the solution produced, and its neutral character when tested with litmus.

We learn from this that sugar is very soluble in water. Careful experiments have shown that water will dissolve about three times its own weight of sugar. This experiment also teaches us that it is neither acid nor alkaline in reaction, but belongs to the class of neutral bodies.

Experiment 231.—Place a small piece of white sugar on an iron plate and heat it gently at first, and then strongly.

Observe that a peculiar odour is given out, and that on further continuing the heating it leaves a residue of carbon.

We learn from this that by heating sugar the dark-coloured body known as **caramel** is produced, and finally, on the further application of heat, carbon only is left.

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Experiment 232.—Heat a little white sugar with water in an evaporating dish. Pour some of the thick liquid out on to a cold stone slab. Now heat that which remains in the dish to a higher temperature.

Observe that the substance melts at about 160° C., and that if allowed to cool it solidifies as a transparent glassy solid. As the heat is continued the body passes into caramel or "burnt sugar."

We learn from this that a non-crystalline mass of **barley sugar** may be prepared by melting sugar with a little water at about 160° C., and that caramel is produced when such a mixture is heated to about 200° C.

Experiment 233.—Pour a little honey into a test-tube, and shake well with an equal volume of alcohol. Pour off the alcohol, and evaporate it to dryness over a water-bath.

Observe that a residue remains in the evaporating dish which is not so sweet as cane sugar. Notice also that the washed honey leaves a residue.

We learn from this that **levulose**, a variety of glucose, may be obtained from honey, and it is more soluble than the grape sugar which remains after washing with alcohol.

SUMMARY TO CHAPTER XXVII.

SUGAR

- **Occurrence in Nature** — Found in most ripe fruits and in other parts of plants, chiefly as fruit sugar or glucose, and cane sugar or sucrose. *Sucrose*, from palm, maple, beet, and sugar-cane. *Glucose*, *dextrose*, or *grape sugar*, found in ripe fruits, and also prepared from starch.
- **Composition** — *Cane sugar* and *grape sugar* consist of the chemical elements *carbon*, *hydrogen*, and *oxygen*. In both varieties hydrogen and oxygen are present in the same proportion as that in which they exist in water, but cane sugar contains more carbon in proportion to the hydrogen and oxygen than grape sugar. All sugars are *carbo-hydrates*.
- **Preparation** — Common sugar is prepared from the beet, sugar-cane, and maple. The vegetable substances are crushed and the saccharine juice is concentrated and then purified by means of albumin or lime, and afterwards colouring matters are removed by the aid of powdered carbon.
- **Properties** — When pure is a white, sweet, neutral, crystalline solid, soluble in water, and is a food-stuff.

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QUESTIONS ON CHAPTER XXVII.

1. How does sugar occur in nature?
2. Describe the preparation of loaf sugar.
3. Say what you know of the origin of molasses.
4. How is sugar candy prepared?
5. What chemical elements are present in sugar
6. How does cane sugar differ from grape sugar
7. Why is sugar a valuable food stuff?
8. What is caramel? How is it prepared? To what uses is it applied?
9. What are carbo-hydrates? Why are they so called?
10. Explain how you would, by means of a few simple experiments, demonstrate the preparation of white sugar.

Preparation for Twenty-eighth Lesson.

APPARATUS AND REAGENTS REQUIRED.

Specimens of different varieties of starch. Microscope and microscopic slides of starches, or good pictures of the same, beakers, evaporating dishes, glass rods, tripod and sand bath, absolute alcohol, litmus papers, litre flask, dilute sulphuric acid, iodine solution, specimens of dextrine, glucose, inuline, gums, woody fibre, cellulose, malt.

EXPERIMENTS TO BE PERFORMED.

Point out the peculiar structure of starch and absence of crystalline form.

Demonstrate that starch does not dissolve in cold water, but on boiling some dissolves.

Show that starch both solid and in solution gives a blue colour when iodine is added to it.

Moisten starch with very dilute hydrochloric acid, and to convert it into a gum, which is thus soluble in water.

Show that dextrine is stained brown by iodine.

Demonstrate that dextrose is not stained by iodine.

Prepare dextrose from starch.

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CHAPTER XXVIII.

STARCH.

A neutral substance composed of Carbon, Hydrogen, and Oxygen—Composition—Peculiar structure; not crystalline—Is found in all parts of a plant—Is obtained from Wheat, Rice, Potatoes, Arrowroot, etc.—Starch in its ordinary condition insoluble in Water—When Starch Powder is boiled with Water, the Membrane of Starch Cells bursts, and the Starch is partially dissolved—Strong Solutions form a Jelly when cold—Used for stiffening Linen—Starch recognized by its forming a blue compound with Iodine—Undergoes no change in the Air at ordinary temperatures; if heated to about 300° F. it becomes slightly discoloured and is changed into a soluble body known as British Gum (Dextrine)—If small amount of Nitric or Hydrochloric Acid be added to the Starch this change is more rapid—Extract of Malt also changes Starch into soluble compounds—Starch as a food.

STARCH is a neutral body which is found in all plants; it is especially abundant in seeds and tubers, and in the stems of the sago-palm, etc. The food value of many vegetables depends upon the quantity of this substance which they contain. It is obtained from wheat, rice, arrowroot, and potato, as a fine white non-crystalline powder which, when examined under the microscope, is found to consist of minute granules, having a peculiar and characteristic structure. This structure is so distinct for the different kinds of starch that it is possible to recognize the origin of the starch by the examination of the granules. The granules of potato starch, which are shown in Fig. 107,

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consist of a number of concentric layers arranged round a nucleus, and they are held together in masses in vegetable cells. Green plants—those containing chlorophyll—manufacture their own starch from the carbon dioxide and water which they take as food.

Starch in its ordinary condition is insoluble in water, but when heated in water to about 70° C. the granules are softened, the membranes burst, and some of the starch dissolves. One of the objects of cooking vegetables is to break up the cells in which the starch granules are stored, to soften, and to a certain extent dissolve out the granules themselves. When starch is boiled with water it forms a solution which sets as a firm jelly on cooling; this substance is largely used in the laundry for stiffening linen.

Fig. 107.—Starch granules from potato (*very highly magnified*).

Starch may easily be recognized by the characteristic blue which it yields with a solution of iodine; this colour disappears on heating, and is therefore not produced when iodine is added to a hot solution of starch, but it reappears on cooling. In the solid state it also assumes a dark blue colour when moistened with iodine. Starch is a very stable body, and it undergoes no change when exposed to air at the ordinary temperatures, but when heated to about 160° C. it passes into a substance which is soluble in water and is known as **dextrine** or **British gum**; the addition of dilute acids, such as hydrochloric or nitric acids, facilitates this change. Prolonged boiling with dilute acids results in the conversion of

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starch into **dextrose**, a variety of glucose. The ferments which are present in malt, saliva and pancreatic juice, all possess the power of converting starch into soluble compounds. We have seen above that starch is insoluble, and in this state it exists locked up in the seeds of plants. During the process of germination it is slowly changed into soluble bodies which can pass into the structure of the young plant, where they serve as nutriment. The vegetable ferment which is developed in the sprouting seed and which brings about this change is known as **diastase.** Advantage is taken of the action of this substance in the preparation of **malt**, for the barley seeds which are rich in starch are placed under such conditions, that they rapidly undergo the early stages of germination, during which the starch is largely converted into a variety of glucose known as **maltose.** When this condition is reached the growing plant is killed by heat. The malt, which is rich in maltose, is afterwards macerated in water and allowed to undergo alcoholic fermentation, by which means the malt sugar is changed into alcohol. Thus then we see that starch is the starting point in the preparation of alcoholic malt liquors.

Only those substances which are soluble in water can pass from the digestive canals of animals into their blood streams ; **starch being insoluble** is changed by certain of the digestive juices into the soluble bodies, maltose and dextrin, before it enters into the blood.

For very young children starchy foods are of little or no value, and in many cases they may be directly harmful, for the simple reason that the ferments which are necessary for their change into soluble substances are not present in

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the digestive juices at this early period of life. Milk, the natural food of a very young infant, contains no starchy matter, but a variety of sugar known as lactose. Nature, therefore, does not provide at this early stage of life any material for the digestion of starch. But we have seen above that a vegetable ferment will produce a similar change in the condition of starch to that which is brought about by the natural digestive juices of animals. Advantage is taken of this fact in the manufacture of the best artificial food for young infants, in which preparation the starch of the farinaceous material employed is wholly converted into such soluble products as can readily pass into the blood.

For commercial purposes starch is chiefly obtained from potato, rice, and maize. When the tubers of the former or the seeds of the latter are crushed and mixed with water and washed by a slow stream, the starch is separated and soon forms a deposit at the bottom of any vessel in which the operation is conducted; while the lighter fragments of pulp, husk, and cellular tissue are carried away in the current. The starch may be afterwards collected and dried.

Experiment 234.—Cut a potato into two pieces, and rub the cut surfaces together in a fine stream of water. Catch the water which washes the cut faces of the vegetable in a beaker, and allow it to stand for some time. Test the powder obtained with red and blue litmus.

Observe that a fine white powder is thrown to the bottom of the beaker. If a small quantity of this white precipitate be mixed with water and heated, it will be seen to thicken at once in the manner which is so characteristic of boiled starch.

We learn from this that potato starch may be obtained by washing the tuber, and that starch is a neutral body.

Experiment 235.—Stir a small quantity of starch up with cold water in an evaporating dish until a thin cream is formed; then add about twice the volume of boiling water, and stir the mixture well.

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Observe that the starch is insoluble in cold water, but that it forms a paste and enters into solution when the dish is heated.

We learn from this that starch may be dissolved to a considerable extent by means of hot water. The temperature most favourable to this change, as we have seen above, is about 70° C. When this operation is carried on upon a slip of glass, so that the starch granules may be observed under a microscope, they are seen to take up water and swell until each granule bursts its enveloping membrane. The contents escape and mainly dissolve in the liquid, whilst the empty sacs float about in the water.

Experiment 236.—Add a few drops of a solution of iodine to a portion of the *cooled* dissolved starch prepared in experiment 234.

Observe the beautiful blue colour which is produced, and how it is discharged by heat, but reappears on cooling the mixture.

Experiment 237.—Place in a litre flask about 400 cc. of a solution of starch, prepared as described in experiment 235, and add to it with constant stirring 3 cc. of dilute sulphuric acid which has been prepared by mixing one part of strong acid with five parts of water. Heat the mixture for some time nearly to boiling point. From time to time withdraw a little of the liquid by means of a glass tube, and test it with iodine solution.

Observe that the iodine becomes less and less active, and ultimately a point is reached at which the iodine solution turns the liquid reddish-brown. When this stage is attained, if a small quantity of the solution be poured into a beaker containing *absolute alcohol*, a flocculent matter separates out which may be dried and tested with iodine.

We learn from this that starch may be converted into dextrine or British gum by the action of heat, and that this substance may be distinguished from starch by its appearance, the brown coloration it yields with a solution of iodine, and its solubility in water.

Experiment 238.—Boil a portion of the solution obtained in the last experiment for fifteen minutes and allow it to cool, and then add to a portion a few drops of iodine solution.

Observe that the substance now gives *no coloration* with iodine, for the starch has passed into **dextrose** or **glucose.**

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We learn from this that insoluble starch when boiled with dilute acids combines with water and passes into, a soluble substance, a variety of sugar known as glucose, the body described in our last chapter.

Inuline is a material which very closely resembles starch in composition and appearance. It occurs in the Jerusalem artichoke, the roots of the dahlia, chicory, etc. It differs from starch in that it dissolves in hot water, and does not form a jelly on cooling. This body is converted by boiling with dilute acids into **levulose,** the variety of sugar which is present in honey along with glucose (see experiment 232).

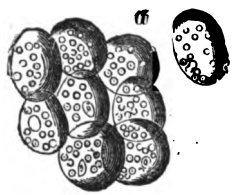

Fig. 108.—Vegetable cells (*highly magnified*).

Gums are also very nearly related both to starch and inuline, for they consist of the same chemical elements in nearly the same proportion; but contain in addition some base such as lime, potash, or magnesia. Gum arabic, which is very extensively employed for industrial purposes, is, for example, a compound of the organic body known as Arabic acid with one of the above-mentioned bases.

Cellulose also belongs to the same class of materials as inuline and starch. It is very largely present in all vegetables, and occurs nearly pure in cotton wool, flax, hemp, etc. The cell walls of plants and all the softer parts contain cellulose, and the hard and woody portions consist also of a hardened variety of cellulose, which has become modified by the addition of more or less resinous matters. This substance is insoluble in water, but when boiled with dilute acids it is slowly converted into the variety of sugar which we have called glucose.

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SUMMARY TO CHAPTER XXVIII.

STARCH

- **Occurrence in Nature** — Found in all plants, largely in seeds of *cereal grains*, tubers of potato, stems of sago palm.
- **Composition** — All the varieties of starch have the same ultimate chemical composition; they consist of the elements carbon, hydrogen, and oxygen, and belong to the class of bodies known as *carbo-hydrates*.
- **Preparation** — By first crushing the vegetable substance and then washing well with a steady stream of water. The powder obtained is afterwards dried.
- **Properties** — White non-crystalline powder, insoluble in cold water, but by the action of hot water it is mainly dissolved and forms a jelly-like mass on cooling. When boiled with dilute acids it is changed into *dextrine*, and this is slowly converted into dextrose.

BODIES ALLIED TO STARCH

- **Inuline** — Found in many plants, but especially in Jerusalem artichoke, dahlia, and chicory.
- **Cellulose** — Found in all plants; nearly pure in cotton wool, flax, hemp, etc.
- **Gum** — Exudations from many trees. Consists of an organic acid body, allied in composition to starch, united with an inorganic base.

QUESTIONS ON CHAPTER XXVIII.

1. How is starch found in nature?
2. Explain how you would obtain a small quantity of potato starch.
3. What chemical elements are present in starch, and why is it called a carbohydrate?
4. How is starch changed by the action of hot water and dilute acids?
5. Describe the relationship which exists between starch, dextrine, and dextrose.
6. What other bodies found in plants are closely allied to starch?
7. Describe the iodine test for starch.
8. How would you prepare dextrose from starch?
9. What are the distinguishing properties of starch?
10. What are starch granules, and where are they found?
11. State what you know of cellulose, inuline, and gum.
12. What variety of sugar may be obtained from inuline?
13. How may starch be converted into sugar?

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Preparation for Twenty-ninth Lesson.

APPARATUS AND REAGENTS REQUIRED.

Specimens of various albuminoids such as casein, fibrin, and albumen of cheese, clot of blood, white of egg. Piece of linen or muslin, large dish, test-tubes, litmus paper, soda-lime.

EXPERIMENTS TO BE PERFORMED.

Tie some flour up in a piece of muslin, and knead it for some time in a basin of water; the starch comes through and will settle to the bottom of the vessel and can be collected and examined; the gluten remains in the bag.

Prove that gluten contains nitrogen by heating with soda-lime and driving off ammonia.

Show that gluten contains carbon, hydrogen, and oxygen by heating in a dry tube and thus obtaining water and a residue of carbon.

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CHAPTER XXIX.

GLUTEN.

If Flour is tied in a Muslin Bag, and well kneaded in a Basin of Water, the Water becomes milky, and on standing Starch sinks to the bottom—All the Starch in the Flour can thus be removed, and then a sticky substance remains in the bag called Gluten—About 70 % of Flour is Starch and 10 % is Gluten—Gluten contains Nitrogen, Starch does not—These Bodies represent two most important Constituents of Food—The Gluten exposed to the Air soon decomposes and smells very disagreeably (putrifies).

GLUTEN is one of the most important forms of **nitrogenous matter** found in plants. It is a sticky, elastic substance which may be obtained from wheaten flour by washing it with water. When exposed to air it decomposes and gives off a disagreeable odour not unlike that of decaying cheese, but when quite dry it does not putrify. This body is a member of a most valuable class of food-stuffs called **albuminoids** or **proteids** ; it is closely allied to *casein of cheese*, *fibrin of blood*, and *albumen of white of egg*, and like these it consists of the chemical elements *nitrogen*, *carbon*, *hydrogen*, and *oxygen*.

Sugar and starch are, on the other hand, representatives of the class of food-stuffs known as **non-nitrogenous.**

The **proteids** possess many properties in common, and have nearly the same composition; they form the essential constituents of all living things. It appears that they

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are *never produced* in animals but only undergo transformation, but *in plants they are actually built up* from simpler bodies.

FLOUR owes its value as a food to the fact that it contains the two most important food-stuffs, starch and gluten; ordinary flour contains about 70 per cent. of the former body, and nearly 10 per cent. of the latter. The gluten may be obtained from flour by loosely tying a small quantity in a piece of muslin or old linen, and kneading it well in a basin of water. The starch is slowly washed out and gives the water a milky appearance; after repeated washings it is almost wholly removed, and may be tested for by means of iodine solution. The substance which remains in the bag is somewhat greyish in colour, and is so sticky and tenacious, that it may be drawn out into long thin elastic threads.

Experiment 239.—Tie some flour up in a muslin bag, and knead it well for some time in a basin of water.

Observe that a white deposit of starch is thrown down from the washing water, and that a sticky substance remains in the bag.

We learn from this that starch and gluten may be obtained from wheaten flour.

Experiment 240.—Dry a portion of the gluten obtained in experiment 239, and mix it well with a small quantity of soda-lime (see experiment 146). Introduce the mixture into a test-tube, and strongly heat it. Place a piece of red litmus paper in the escaping gas.

Observe that the gas which is given off turns the litmus paper blue.

We have learnt in experiment 142 that the *volatile alkali is ammonia*, and in experiment 146 that organic bodies which give off ammonia, on heating, contain nitrogen; since gluten yields ammonia we may infer that it contains nitrogen, and belongs therefore to the class of nitrogenous bodies.

Experiment 241.—Strongly heat another portion of the dried gluten obtained from experiment 239 in a hard glass tube.

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Observe that water is given off, dense fumes escape, and after a time a residue of carbon remains.

We learn from this that gluten contains carbon; since it yields water we may infer that it contains hydrogen and oxygen. Therefore we may conclude that it contains in all the four elements—nitrogen, carbon, oxygen, and hydrogen.

Experiment 242.—Expose a piece of moist gluten to the air.

Observe that after a few days it becomes covered with mildew, and emits a very unpleasant odour.

We learn from this that gluten, like other nitrogenous organic bodies, will readily putrify when exposed to air.

As mentioned above gluten is a most valuable food-stuff. By the process of digestion this and other albuminoid bodies are rendered soluble, and in this condition they are capable of being absorbed by the blood.

SUMMARY TO CHAPTER XXIX.

GLUTEN

- **Occurrence in Nature** – This body is found in the grain of wheat, and is therefore present in flour, bread, etc. It is closely allied to a number of substances found in other plants, such as *legumen* and vegetable *albumen*.
- **Composition** – Contains the chemical elements *nitrogen, hydrogen, carbon,* and *oxygen*. Gluten is included in the class of *nitrogenous organic bodies* known as proteids, or albuminoids. These bodies serve in the food of animals as tissue formers, or structure builders. They are all of very complex composition, but the result of many analyses shows that they vary in composition according to the following table :—

Carbon	52·7 to 54·5	per	cent.
Hydrogen	6·9 „ 7·3	„	„
Nitrogen	15·1 „ 16·5	„	„
Oxygen	20·9 „ 23·4	„	„
Sulphur	·8 „ 1·6	„	„

- **Preparation** – Gluten may be obtained, nearly pure, from wheaten flour, by washing out the starch.
- **Properties** – It is a sticky, elastic body, which soon decomposes when exposed to moist air. It is one of the most valuable constituents of foods.

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QUESTIONS ON CHAPTER XXIX.

1. Where is gluten found in nature?
2. How would you obtain a specimen of gluten?
3. How would you prove that gluten contains nitrogen?
4. Why is gluten a valuable food stuff?
5. Name other bodies which belong to the same class of substances as gluten.
6. How would you prove that gluten contains carbon?
7. What are the chief properties of gluten?
8. How would you distinguish gluten from starch?
9. What chemical elements are present in gluten?
10. What changes does gluten undergo in digestion?

Preparation for Thirtieth Lesson.

APPARATUS AND REAGENTS REQUIRED.

Specimens of absolute alcohol, fusel oil, methylated spirit, wood spirit, proof spirit, glucose, dried yeast, milk fresh and sour. Apparatus as in Fig. 109, flasks, retort and condenser, test-tubes, watch-glass, beaker, short test-tube with cork and delivery tube drawn out to a jet.

EXPERIMENTS TO BE PERFORMED.

"Show that alcohol is neutral liquid dissolving in water, that it burns with nearly colourless flame, and leaves no residue of carbon.

Show that it can be made to boil at a much lower temperature than water by placing a test-tube of it in hot water.

Distil beer, and collect the alcohol and water which come over; add quicklime to this. Allow it to stand some hours and distil again. Show that this is much stronger, catches fire readily, and has a more burning taste.

Make a solution of common sugar in a large flask; add yeast, and fit a cork with a bent tube to the flask. Let the tube dip into lime-water.

Place it in a warm place, and after some days show that spirit has been formed in the flask by distilling the liquid and collecting the portion coming over first."—*Science Department's Syllabus.*

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CHAPTER XXX.

SPIRIT.

Alcohol, Spirits of Wine—A colourless, light Liquid—Neutral to Test Papers—Has pleasant odour, boils at 173° F.—Burns with a flame, which gives very little light, without leaving any black residue of Carbon.—A large number of different bodies dissolve in it—It is the intoxicating principle in Wines and Spirits—In Beer there is 3 to 5 per cent. of Alcohol—In Light Wines about 8 per cent.—In Spirits 50 to 75 per cent.—The different flavours of Wines and Spirits depend on very small quantities of other bodies present—Alcohol dissolves in Water, giving out Heat—"Proof Spirit" contains 50·76 parts of Water and 49·24 parts of Alcohol—If more Water be present the Spivit will not set fire to Gunpowder when burning—Alcohol obtained from Grape Sugar—Fermentation—Grape Sugar converted into Alcohol and Carbon Dioxide by presence of some ferment such as Yeast—Cane Sugar on the addition of Yeast is first converted into Grape Sugar, then into Alcohol and Carbon Dioxide—Use of Yeast in Brewing—Yeast not necessary for making Wine as there is already a ferment in expressed juice of Grape.

ALCOHOL is the name by which we know one particular representative of a great class of substances called **the alcohols.** The term is *commonly* used for the body which is present in **spirits of wine,** and which is the active principle of most alcoholic liquors. We should remember that the name is, in the next pages, to be used only for the best known member of a great group of substances.

When pure, alcohol is a clear, colourless, neutral liquid, which has a slightly pleasant odour and strong burning taste. It will burn in air with a flame which gives out much heat but very little light, and it burns without leaving

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any black residue of carbon. Alcohol is lighter than water, but being soluble in that body it mixes with it freely in all proportions. With other substances it exists in wines, spirits, and beer, of which it constitutes the intoxicating agent. The different flavours of wines and spirits depend upon the presence of small quantities of other flavouring and odorous substances. It will not only dissolve in water, but a very large number of bodies will dissolve in it; and, in consequence of these facts, many liquids are placed upon the market as wines and spirits, which are innocent, save in name, of the grape and other substances from which they are supposed to be derived. Alcohol may be cheaply distilled from many vegetable fermenting substances; among other sources the potato is one of the most economical. For this reason, potato spirit and water, coloured with caramel and flavoured with extract of French plums, sometimes finds its way into the market as brandy.

FERMENTATION. When a solution containing sugar and a nitrogenous body is exposed to the air at a temperature between 75° and 90° F., what is called fermentation takes place. This fermentation is due to the growth, in the liquid, of a lowly form of life, allied to the fungi. There are several fungoid forms which will live in sugar solutions, and the two most important kinds of fermentation to which they give rise are known respectively as **vinous** or **alcoholic**, and **lactic**. The fermentation may, of course, be prevented by keeping away the living **spores** or **germs** of the ferment, or by making the conditions of the liquid unfavourable to their growth. This latter end may be attained by the presence of bodies which act as poisons to

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the living ferment,—these are called **antiseptics**; or by converting the liquid into a syrup by the presence of too much sugar. The temperature of the liquid will also determine its favourability to the growth of the lowly organisms.

We saw in Chapter XXII. that the fermentation of alcoholic liquids, and their conversion into vinegar by **acetous fermentation,** was due also to the life of a minute and simple organism which was able during its growth to convert the alcohol into acetic acid.

FERMENTS include a number of bodies which give rise to more or less complex chemical changes ; they are sometimes *lifeless chemical products found in living bodies* ; the changes which take place during the conversion of starch into sugar, in malt, and during digestion are brought about by such substances. We have, then, good examples of non-living ferments in the **diastase** of malt and the **ptyâlin** of saliva. At other times the cause of fermentation is, as stated above, *a lowly living organism*, which, during its rapid development, growth, and multiplication *brings about complex chemical changes.* Under this head we should include the **mycoderma aceti** and the **yeast plant.**

The ferment which produces alcoholic fermentation is the torula or yeast plant. This minute organism is present in the dust of the air. The cultivated variety grows in beer and wine. The temperature most favourable to its growth is between 25° and 30° C. As it grows it breaks up the sugar into alcohol and carbon dioxide ; the latter escaping in bubbles from the surface of the liquid gives rise to a **frothing** of the mass, whence the name fermentation. The process continues until sufficient alcohol is produced to

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kill the lowly plant. Solutions of glucose or grape sugar are most favourable to the growth of yeast, but cane sugar will also undergo alcoholic fermentation. Cane sugar, on the addition of yeast, is first converted into grape sugar, and then into alcohol and carbon dioxide. The change which apple juice undergoes in forming **cider** well illustrates the action of this ferment. The juice of the apple contains sucrose, the germs of yeast falling from the air into the unfermented liquid cause it to "**work**"; thus sucrose is changed into glucose, and this is afterwards broken up into alcohol and carbon dioxide; the juice becomes alcoholic and is thus converted into cider.

We saw in Chapter XXVIII. that starch may easily be converted into glucose; the brewer therefore employs malt, barley, and rice in place of sugar. The conversion is brought about by mashing together the bruised grains and a small quantity of malt in water, at about 70° C. By this means the starch is converted, by the non-living ferment **diastase**, into sugar. The saccharine liquid thus obtained is known as "**wort.**" In this liquid yeast is placed, and by its growth it gives rise to **alcoholic fermentation.**

Experiment 243.—Prepare a solution of glucose in water, or neutralise the free acid in the glucose solution prepared in experiment 238 by the careful addition of caustic soda solution. Place the mixture in a flask, and add a small quantity of "**barm**" obtained from the brewer, or dried yeast. Stand the flask in a vessel of warm water, kept at about 25° C. After the liquid has become somewhat turbid and frothy, fit up the flask with a cork and delivery tube, and allow the tube to dip under warm water in a shallow pneumatic trough. Invert a jar full of warm water over the end of the tube, as shown in Fig. 109. Keep the apparatus in a warm place for several days.

Observe that after a time a gas collects in the flask which will extinguish a burning taper, and which answers to the other tests for carbon

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dioxide. Notice also that the liquid in the flask becomes turbid, and that when a small quantity is boiled it gives off a decided spirituous odour.

We learn from this that by the growth of yeast a sweet solution of sugar is rendered alcoholic or spirituous and carbon dioxide is liberated.

Experiment 244.—After several days pour the contents of the flask (Fig. 109, experiment 243) into a retort and distil over about one-third of the whole by means of the apparatus shown in Fig. 110.

Observe that the distillate has an unpleasant, but spirituous, odour, and that it burns with an almost colourless flame which gives out intense heat.

We learn from this that alcohol may be obtained from the fermented saccharine liquid by distillation. The unpleasant odour is due to a body known as **fusel oil**; this is a mixture of alcohols with higher boiling points than that of common alcohol.

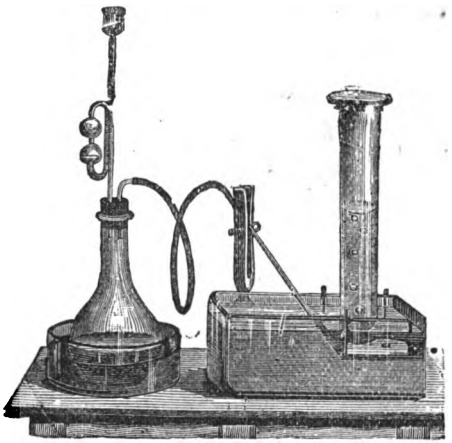

Fig. 109.—Liberation of carbon dioxide during alcoholic fermentation.

Experiment 245.—Introduce into a retort *k*, about half a pint of old ale; place the bulb of the retort in a water bath *b*, and heat the same. Condense the vapour which comes over in a flask *v*, which should be kept cool by means of a gentle stream of cold water.

Observe that a clear colourless mobile liquid collects in the condenser; this should be cooled, and then mixed with quicklime and allowed to stand for a day or two; on re-distilling, strong alcohol may be obtained.

We learn from this that alcohol may be separated from malt liquor by distillation. Since water vapour comes over with the alcohol, the distillate is only dilute alcohol. As

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water will combine with lime and the alcohol will not, the liquid may be freed from water to a considerable extent by redistillation with lime. By repeating this operation several times **absolute alcohol** may be obtained.

Alcohol thus obtained proves on examination to be a clear, colourless, neutral liquid which is lighter than water, having a specific gravity of ·794 and boiling at 78° C. It exists in variable proportions in alcoholic drinks, beer contains 3 to 5 per cent., light wines such as claret about 8 per

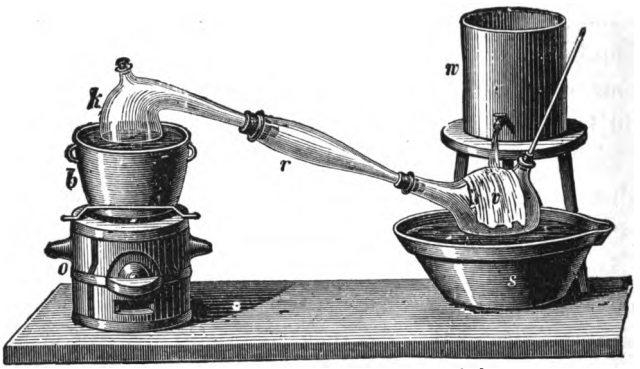

Fig. 110.—Preparation of alcohol from old ale.

cent., stronger wines like sherry from 14 to 18 per cent., and spirits such as gin, rum, whisky, and brandy, from 50 to 75 per cent. The different appearances are due to the presence of small quantities of colouring matter, and the characteristic odours and flavours to essential oils. Alcohol dissolves in water giving out heat; it removes water from animal tissues with which it is brought in contact, and hardens the albumen, and is therefore used in the preservation of anatomical specimens and in hardening histological preparations. The same chemical elements as are present

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in starch, sugar, cellulose, etc., make up alcohol, for it is composed of 52·2 per cent. of carbon, 34·8 per cent. of oxygen, and 13 per cent. of hydrogen.

Spirits of wine is a mixture of alcohol with a small quantity of water, and, strictly speaking, the name should only be employed for the spirit obtained by distilling wines; but, as we have seen, alcohol may be obtained from other liquids; hence the name has been somewhat loosely employed. From spirits of wine absolute alcohol may be obtained by the method described in experiment 245.

The Government duty on absolute alcohol is very high; but an impure variety known as **methylated spirit**, containing 90 per cent. by volume of strong spirit, and 10 per cent. of *wood spirit* or *methyl alcohol*, is sold for chemical and manufacturing purposes. This substance has an exceedingly unpleasant odour, and is useless for drinking purposes; the Excise therefore permits it to be sold duty free.

The so-called **proof spirit** of the Excise Department contains 49·24 per cent. of alcohol, and 50·76 per cent. of water. When gunpowder is moistened with this liquid and a light applied, the whole is ignited, but when more water is present the diluted spirit will not set fire to gunpowder when burning.

Experiment 246.—Pour a small quantity of absolute alcohol into a small beaker; test it with red and blue litmus papers, then add water.

Observe that the alcohol has no action upon the litmus, and that it readily dissolves in water in all proportions.

We learn from this that alcohol is a neutral body, and it is soluble at once in water.

Experiment 247.—Place a small quantity of alcohol in a watch-glass, apply a light, and draw a piece of cold porcelain through the flame.

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Observe that the liquid burns with an almost colourless flame; it leaves no residue, and does not deposit carbon upon the cold body.

We learn from this that alcohol burns with a flame which gives little light but very great heat, and does not leave a black residue of carbon.

Experiment 248.—Fit up a short test-tube with a cork and delivery tube drawn out to a point. Introduce into the tube some absolute alcohol, and plunge the apparatus into *hot, but not boiling, water*.

Observe that the liquid boils, and that a vapour escapes which may be ignited at the jet, where it burns with the characteristic alcohol flame.

We learn from this that alcohol boils at a lower temperature than water; for this reason we heated the ale in experiment 245 over a water bath, for by keeping the contents of the retort just below the boiling point of water, we were able to drive off the more volatile alcohol without sending over with it much water vapour.

Experiment 249.—Pour a small quantity of alcohol into a test-tube and introduce a piece of shellac, and shake the mixture well.

Observe that the substance dissolves at once.

We learn from this that alcohol will dissolve shellac and form common "**spirit varnish.**" Such substances as camphor, iodine, oils, and resins will dissolve in alcohol.

The manufacture of spirits on a large scale involves—

(*a*) The *conversion of starch into sugar* by the non-living ferment *diastase*.

(*b*) The *fermentation of sugar into alcohol* by the action of the living ferment, *yeast*.

(*c*) The *separation of the spirit* by the operation of distillation.

The alcohol present in **wines** made from grape juice is derived from the fermentation of the glucose, which is present in the ripe fruit; the ferment, which is a body analogous to yeast, is derived from the skins of the grape. The spirit distilled from wine is known as **brandy**, and this is strictly spirit of wine; it owes its flavour when thus prepared to an essential oil derived from the grape.

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Lactic fermentation changes lactose or milk sugar into **lactic acid** during souring, but milk has been kept sweet for years by boiling and excluding all germs.

Putrefaction is the variety of fermentation which is set up in nitrogenous organic bodies, which results in the liberation of ill-smelling substances.

Infections diseases are also due to the growth in the blood and other tissues of the body of minute living organisms, which give rise to a form of fermentation.

SUMMARY TO CHAPTER XXX.

ALCOHOL

- **Composition** – Chemical compound, consisting of the elements carbon, hydrogen, and oxygen.
- **Preparation** – Produced on a large scale by the natural fermentation of saccharine solutions. This action known as *vinous fermentation*. From alcoholic liquids it is obtained by distillation.
- **Properties** – Colourless, neutral, volatile, combustible liquid; lighter than water, in which it dissolves giving out much heat. When it burns in air, water and carbon dioxide are produced; it is the intoxicating principle of alcoholic liquors.
- **Kinds** – Common alcohol is a representative of a great class of bodies, to which *wood spirit*, *fusel oil*, etc., belong,—*methylated spirit*, an impure variety of alcohol, containing *methyl alcohol*. *Proof spirit* contains 49·2 per cent. of alcohol. *Spirits of wine* is, strictly speaking, the name of alcohol distilled from wine.

QUESTIONS ON CHAPTER XXX.

1. How would you illustrate the chief properties of alcohol?
2. What do you understand by fermentation? Give examples.
3. Name three ferments, and say what they do.
4. Explain how you would demonstrate the fermentation of saccharine solutions.
5. What changes take place when sugar undergoes fermentation?
6. What is the name of the ferment which gives rise to vinous fermentation, and what changes does it bring about in cane sugar?
7. How would you obtain alcohol from old ale?
8. What is absolute alcohol? How is it prepared?
9. State the proportion of alcohol in common alcoholic drinks.
10. For what industrial purpose is alcohol employed?

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Specimen Examination Papers

ISSUED BY THE DEPARTMENT OF SCIENCE AND ART.

The alternative syllabus for Stage I. of Inorganic Chemistry is intended for those students who only require the Elements of Chemistry as a foundation for their studies in other subjects.

A paper will be set in each syllabus at the examination in May, and a candidate may take either, but not both. His success will be registered as a success in the elementary stage of Inorganic Chemistry, and no second payment will be made on account of the same class at a future examination in that subject.

ALTERNATIVE FIRST STAGE OR ELEMENTARY EXAMINATION, 1887.

You are permitted to attempt eight Questions.

21. A glass of water is exposed to the air. In time the water disappears into the air. How would you account for this? How could you prove that there is moisture in the air?
22. Air is passed over red hot iron. What change does this cause in the air and in the iron?
23. How would you show that the gas obtained by dissolving marble in hydrochloric acid is also obtained in the breath?
24. Two samples of water are given to you. One is hard water and the other is distilled water. Describe two methods of distinguishing between them.
25. What is vinegar? How is it prepared? Vinegar is poured upon washing soda. What happens?
26. Ammonia is classed as an alkali. Why? Name some of the sources from which it can be obtained, and give its composition.
27. A piece of lead, a piece of copper, and some mercury, are separately heated in a crucible over a lamp. Describe what occurs in each case.
28. From what substances can starch be obtained? Of what is it composed, and how does it behave when boiled with water?
29. Name some commonly occurring compounds of sodium. How can you show that chlorine is a constituent of common salt?
30. What substances are contained in flour? How can they be separated, and what essential difference is there in their composition?
31. What is meant by that a solution is saturated? How would you prove that no loss of weight occurs when a substance is dissolved in water?
32. What are the distinguishing characters of cast iron, wrought iron, and steel? What is iron rust

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MAY, 1888.

21. Some soil is shaken up with distilled water. How could you prove that some of it has been dissolved?
22. Describe an experiment to show that when a substance burns in air, oxygen combines with the burning body.
23. What is the chemical nature of fat? How may it be converted into soap? What is the difference in chemical composition between "hard" and "soft" soap? Describe the action of soap upon hard water.
24. What happens when sulphur burns in air? How would you account for the peculiar smell of burning sulphur?

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25. What is the composition of limestone? Name other substances of the same composition. Explain what occurs when limestone (1) is strongly heated and (2) is placed in dilute hydrochloric acid.
26. How would you prove that hot water dissolves more nitre than cold water?
27. Of what chemical elements is sugar composed? What happens when a solution of sugar is mixed with yeast?
28. What is the chemical nature of clay, gypsum (plaster of Paris), glass, and chalk?
29. What is the principal source from which tartaric acid is obtained? Of what elements is it composed? Describe what happens when it is added to a solution of carbonate of soda?
30. In what respects does metallic lead differ from red lead? How can the one be converted into the other?
31. How can you obtain alcohol from beer? What is proof spirit?
32. Of what metals is a penny composed? What is the difference between an alloy and an amalgam?

MAY, 1889.

21. Describe an experiment to show that air has weight.
22. How would you prove by experiment that air, water, and iron expand on being heated?
23. Why is the gas obtained by the action of hydrochloric acid upon the black oxide of manganese termed "Chlorine"? Give some account of the properties of this gas, and name any substances known to you that contain it.
24. A piece of charcoal is heated in the air until it disappears, leaving only a white residue (ash). Give an explanation of this change.
25. What is the effect of boiling a solution of sal-ammoniac with lime or potash? Name the gas involved, and state from what other sources it can be obtained.
26. Group the following bodies as elements or compounds, and divide the elements into metals and non-metals: lead, brass, brimstone, starch, graphite, mercury, glycerine, clay, alum, tin, calcium.
27. Name the common ores of iron, and state how cast iron is obtained from them.
28. What chemical elements are found in chalk, common salt, washing soda, blue vitriol, brass, sugar of lead, sugar, and soap?
29. How is cane sugar obtained, and in what respects does it differ from grape sugar?
30. What is gluten, and how is it obtained?
31. How may acetic acid be prepared (1) from alcohol, and (2) from wood? In what respects does acetic acid resemble and differ from sulphuric acid?
32. What is the chemical nature of tallow, butter, soap, oil?

MAY, 1890.

21. How can you show that the volume of a gas is changed by temperature or pressure?
22. In what respects do hard, soft, mineral, and distilled waters differ from one another? In which class do you place Thames water? How would you distinguish a mineral water from a distilled water?
23. Describe an experiment to prove that a substance when burnt in air gains in weight. How do you account for this increase in weight?
24. When charcoal is burnt in air or in oxygen heat is involved. How is the amount of this heat expressed?
25. Turpentine put into chlorine gives rise to a puff of black smoke, phosphorus put into chlorine takes fire, and moist red cloth put into chlorine loses its colour. Explain these changes.
26. What are the most characteristic properties of metallic lead? How can you prove that pure water can dissolve a small amount of it?
27. What is the chemical nature of soap, sugar, starch, and fat?
28. Why is mercury called a metal? What are its properties? Why is it used in the construction of barometers and thermometers?
29. What is iron rust? How is it formed? A piece of iron is put into dilute hydrochloric acid. What happens?
30. What is tallow? What are its properties? Into what substances can it be decomposed?
31. How is vinegar obtained? To what substance is its acidity due? In what respects does this acid substance differ from sulphuric acid, and in what respects does it resemble it? How could you obtain sugar of lead from vinegar?
32. The following bodies are placed on a piece of sheet iron, which is gradually made red hot: tartaric acid, mercury, starch, chloride of ammonium, and sulphur. What are the changes which each substance will undergo?

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MAY, 1891.

21. State any facts known to you which serve to show that in the case of substances which are soluble in water, more of a solid is dissolved by hot water than by cold water, and more of a gas is dissolved by cold water than by hot water.

22. Why does a combustible substance burn more brilliantly in oxygen than in air? Give some account of the chemical changes which occur when we burn (1) a tallow candle; (2) paraffin oil; (3) a piece of wood; (4) a piece of brimstone, in the air.

23. How could you show that the gas contained in human breath, which renders lime-water turbid, is also contained in limestone, marble, oyster-shells, and chalk?

24. How could you extract "spirit of wine" from beer, and by what tests could you show that the liquid so obtained consists of a mixture of alcohol and water?

25. What is rock salt? How is it converted into ordinary table salt?

Name the metal contained in common salt and describe its properties.

26. What chemical elements are contained in the following substances: hæmatite, brass, galena, clay, plaster of Paris, marble, sugar of lead, mutton suet, starch, soap, and sugar?

27. What is the chemical nature of a fat? Into what substances is tallow resolved on heating it with high-pressure steam? How is a fat converted into a soap? How do you account for the occurrence of the white scum which is found when ordinary washing soap is dissolved in a hard water?

28. What are the two chief substances into which wheaten flour is resolved by treatment with water? Describe the chief properties of each?

29. In what respect does the sugar obtained from honey differ from common sugar? How would you distinguish them by chemical tests?

30. How is quicklime made? What is the chemical change which occurs when lime is slacked? How do you account for the heat evolved?

31. How could you show that when charcoal burns in air it is the oxygen and not the nitrogen that sustains the combustion?

32. Describe an apparatus for making pure drinkable water from sea-water.

MAY, 1892.

21. How may it be shown that water is a *chemical compound* of hydrogen and oxygen, and that air is a *mixture* of oxygen and nitrogen?

22. In what respects do rain-water, river-water, and sea-water differ from each other?

23. Describe the ordinary method of making hydrogen gas. What precaution is necessary to obtain the gas free from atmospheric air?

24. By what simple tests or experiments could you distinguish and identify nitric, hydrochloric, and sulphuric acids?

25. How could you ascertain that spring water contains dissolved gas? By what experiments could you determine the nature of this gas?

26. Metallic copper heated in the air becomes covered with a black scale; lead, similarly heated, becomes first yellow and then red. How could you prove that these products contain oxygen?

27. How is cast iron made from clay-iron stone?

28. How do you prove the presence of carbon in sugar, starch, alcohol, vinegar?

29. What is the use of yeast in brewing? What are the chief products of its action?

30. Classify the following substances as solids and liquids:—

Manganese dioxide, calcium chloride, ammonium chloride, phosphorus, graphite, plaster of Paris, galena, argol, acetic acid, cinnabar, silica, stearic acid, glycerin, gluten, tartaric acid.

31. What chemical elements are contained in plaster of Paris, clay, iron-slag, litharge, red lead, "proof spirit," washing soda, tallow, common salt, bleaching powder, brimstone, ammonia, oil of turpentine?

32. Describe the preparation, properties, and uses of bleaching powder.

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APPENDIX.

SYMBOLS AND FORMULÆ.

Dalton's atomic theory assumes that all elementary matter consists of exceedingly minute, ultimate, indivisible particles called atoms (Gr. *a*, not ; *temno*, I cut), which, with few exceptions, are the same size in different kinds of matter, but of different weights. These ultimate atoms of any one element, we believe, differ in weight and properties from those of all other elements, although all the atoms of any one body have a definite weight. In some cases the differences are very great, in others only slight. These atoms combine together in various proportions, so that when chemical union takes place it is due to the action of atom on atom, or groups of atoms on groups of atoms. Thus, then, if it be granted that elementary matter is made up of atoms, and that these have a fixed weight for any one body, and that they cannot be divided, it must follow that the proportions in which one element combines with another is the same as the number which represents its atomic weight, or some multiple of this number.

Atomic Weight.—The lightest known body is hydrogen; the atom of hydrogen is adopted, therefore, as the unit of comparison, and the atomic weight of any element is the number which represents the weight of one atom of the body

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as compared with that of an atom of hydrogen. Although a description of the methods employed in determining the atomic weight of a body would be beyond the scope of this work, yet the student may, at this stage even, learn that in making such determinations chemists take the following facts into consideration.

(*a*) The smallest proportion by weight of the element which will enter into combination with, or take the place of, one part by weight of hydrogen.

(*b*) The weight of the element which, in the state of a gas or vapour, occupies the same space as one part by weight of hydrogen under like conditions of temperature and pressure.

A molecule is the smallest portion of matter, elementary or compound, which can exist alone. All molecules are built up of atoms. The molecules of the elements generally consist of two atoms. The molecules of compound substances may consist of any number greater than two atoms. For example, the hydrochloric acid molecule consists of two atoms, water of three, ammonia of four, and many organic compounds of a great number.

The various chemical elements are represented by means of certain signs which are called chemical symbols.

Symbols are the characters used to represent atoms of **elements.** The symbol of a body is usually the first letter of its English or Latin name : thus, O stands for one atom of oxygen ; K, for the Latin word *kalium*, stands for one atom of potassium.

In those cases where the names of several elements commence with the same letter, the first letter is used for

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the most important, and for the others some prominent letter is added to the first. For example, the names of many elements commence with **C**, and this single letter is employed to stand for the element carbon, whilst—

Cl	is the	symbol for	Chlorine.	**Cu**	is the	symbol for	Copper (cuprum).
Cr	„	„	Chromium.	**Cd**	„	„	Cadmium.
Ca	„	„	Calcium.	**Ce**	„	„	Cerium.
Cs	„	„	Cæsium.	**Co**	„	„	Cobalt.

It happens that in some cases this plan is impracticable; we then make use of the distinctive letters of the Latin names, thus, in the list given below, see silver and sodium—

Sulphur	is represented	by	**S.**	Scandium	is represented	by	**Sc.**
Silicon	„	„	**Si.**	Silver (*Argentum*)	„		**Ag.**
Selenium	„	„	**Se.**	Sodium (*Natrium*)	„		**Na.**
Strontium	„	„	**Sr.**				

Many of the commoner metals are represented by the more important letters of their Latin names, thus—

Lead	*Plumbum*	**Pb.**	Tin	*Stannum*	**Sn.**
Mercury	*Hydrargyrum*	**Hg.**	Antimony	*Stibium*	**Sb.**
Copper	*Cuprum*	**Cu.**	Gold	*Aurum*	**Au.**
Iron	*Ferrum*	**Fe.**	Silver	*Argentum*	**Ag.**

A chemical symbol not only denotes to the chemist the particular element and one atom of the body, but also one atomic weight's worth, thus—

H	the	symbol for	Hydrogen	means	one atom	and	1	part	by weight.
Cl	„	„	Chlorine	„	„		35·5	parts	by „
O	„	„	Oxygen	„	„		16	„	„ „
S	„	„	Sulphur	„	„		32	„	„ „

This, because one atom of chlorine weighs thirty-five and a half times as much as an atom of hydrogen, and the atoms of oxygen and sulphur weigh respectively sixteen

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and thirty-two times as much as an atom of hydrogen under like conditions. Of the gaseous elements the symbol also denotes one volume ; thus, then—

H	stands for	*one atom, one volume, and*	1	*part by weight of*	*Hydrogen,*
O	„	„	16	*parts by* „	*Oxygen.*
Cl	„	„	35·5	„ „ „	*Chlorine.*
N	„	„	14	„ „ „	*Nitrogen.*

Formulæ are used by chemists to represent the kinds of **atoms present in the molecules**. Thus HCl is the formula for hydrochloric acid, and it implies that one molecule of that body contains one atom of hydrogen and one atom of chlorine.

The molecules of the elements usually contain two atoms, and they are represented, for example, thus—

$$H_2,\ O_2,\ Cl_2,\ N_2,\ Br_2,\ I_2,\ Si_2,\ Fe_2.$$

A number placed at the right hand lower corner of a symbol shows how many atoms there are of that element present in the molecule of the body.

Binary Compounds.—The least complex chemical compounds are those which are composed of two elements only, and are therefore termed binary compounds ; the formulæ for their molecules contain therefore only two symbols, thus NaCl is the formula for the binary compound common salt, or sodic chloride ; again, CaO stands for lime, or calcic oxide, and H_2O represents the molecule of water.

An exponent is a figure always written *below* the symbol of an element to indicate the number of atoms of that element. Thus H_2O is the formula for a molecule of water, and it indicates two atoms of hydrogen and one of oxygen.

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A coefficient is a number sometimes placed before the formula for a compound body to indicate the number of molecules and the number of atoms present. Thus $2H_2O$ means two molecules of water, and, therefore, four atoms of hydrogen and two atoms of oxygen.

Chemical equations are employed to represent chemical changes; the symbols or formulæ of the bodies employed are placed on the left of the sign = which implies here, "yields"; and the resulting substances are represented by formulæ and symbols placed on the right; the sign + is employed in the sense of "acted upon by" or "reacting with," thus—

$Na + H_2O = NaHO + H$, means that one atom of sodium acting upon one molecule of water yields one molecule of sodic hydrate and one atom of hydrogen,

and $H_2 + O = H_2O$,

implies that one molecule, or 2 parts by weight, of hydrogen and one atom, or 16 parts by weight, of oxygen form one molecule, or 18 parts by weight, of water,

and $Mg + O = MgO$,

one atom, or 24 parts by weight, of magnesium combines with one atom, or 16 parts by weight, of oxygen to form 40 parts by weight of magnesic oxide (magnesia).

$$Zn + H_2SO_4 = ZnSO_4 + H_2.$$

This equation informs us that one atom of zinc will react with one molecule of sulphuric acid, forming one molecule of zincic sulphate, and liberating one molecule or two atoms of hydrogen.

For a table of the symbols of the chemical elements, see pages 314, 315.

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TABLE OF THE ELEMENTS WITH THEIR ATOMIC WEIGHTS, SYMBOLS, AND ATOMICITIES.

[NOTE.—*The Non-Metals are Printed in Capitals.*]

Symbol.	Name of Element.	Atomicity.	Accurate Atomic Weight.	Approximate Atomic Weight.
Al	Aluminium . . .	IV.	27·02	27
Sb	Antimony . . .	III. V.	120·00	120
As	ARSENIC . . .	III. V.	74·90	75
Ba	Barium	II.	136·80	137
Be	Beryllium . . .	II.	9·10	9
Bi	Bismuth	III. V.	208·00	208
B	BORON	III.	10·90	11
Br	BROMINE . . .	I.	79·75	80
Cd	Cadmium. . . .	II.	111·90	112
Cs	Cæsium	I.	132·50	133
Ca	Calcium	II.	39·90	40
C	CARBON . . .	IV.	11·97	12
Ce	Cerium . . .	IV.	141·00	141
Cl	CHLORINE . .	I.	35·37	35·5
Cr	Chromium . . .	VI.	52·40	52·5
Co	Cobalt.	II. IV.	59·00	59
Cu	Copper	II.	63·40	63
Di	Didymium . . .	III.	144·00	—
Er	Erbium	III.	166·00	—
F	FLUORINE . .	I.	19·10	19
Ga	Gallium	IV.	69·00	70
Au	Gold	I. III.	196·00	196
H	HYDROGEN . .	I.	1·00	1
In	Indium	III.	113·40	113·5
I	IODINE. . . .	I.	126·53	126·5
Ir	Iridium	IV.	194·00	194
Fe	Iron	II. IV. VI.	55·90	56
La	Lanthanum . . .	IV.	138·50	—
Pb	Lead	II.	206·40	206·5
Li	Lithium	I.	7·01	7
Mg	Magnesium . . .	II.	23·90	24
Mn	Manganese . . .	II. IV. VI.	55·00	55
Hg	Mercury	II.	199·80	200

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TABLE OF THE ELEMENTS WITH THEIR ATOMIC WEIGHTS, SYMBOLS, AND ATOMICITIES—*Continued.*

[NOTE.—*The Non-Metals are Printed in Capitals.*]

Symbol.	Name of Element.	Atomicity.	Accurate Atomic Weight.	Approximate Atomic Weight.
Mo	Molybdenum . .	VI.	95·80	96
Ni	Nickel.	II. IV.	58·60	59
Nb	Niobium	V.	94·00	94
N	NITROGEN . .	III. V.	14·01	14
Os	Osmium	VI.	198·60	199
O	OXYGEN . . .	II.	15·96	16
Pd	Palladium . . .	II.	106·20	106
P	PHOSPHORUS .	III. V.	30·96	31
Pt	Platinum. . . .	IV.	196·70	196·7
K	Potassium . . .	I.	39·04	39·1
Rh	Rhodium. . . .	III.	104·00	104
Rb	Rubidium . . .	I.	85·20	85
Ru	Ruthenium . . .	VI.	104·50	104·5
Sc	Scandium . . .	—	44·00	—
Se	SELENIUM . .	II. IV. VI.	78·80	79
Ag	Silver	I.	107·66	108
Si	SILICON . . .	IV.	28·30	28
Na	Sodium	I.	22·99	23
Sr	Strontium . . .	II.	87·20	87
S	SULPHUR. . .	II.	31·98	32
Ta	Tantalum . . .	V.	182·00	182
Te	TELLURIUM. .	II.	125·00	125
Tr	Terbium	II.	148·50	148·5
Tl	Thallium. . . .	I.	203·60	204
Th	Thorium	IV.	232·40	232·4
Sn	Tin.	II. IV.	117·80	118
Ti	Titanium. . . .	II. IV.	48·00	48
W	Tugsten	IV. VI.	183·60	184
U	Uranium. . . .	II. IV. VI.	240·00	240
V	Vanadium . . .	V.	51·20	51·2
Y	Yttrium	III.	89·50	—
Zn	Zinc	II.	64·90	65
Zr	Zirconium . . .	IV.	90·00	90

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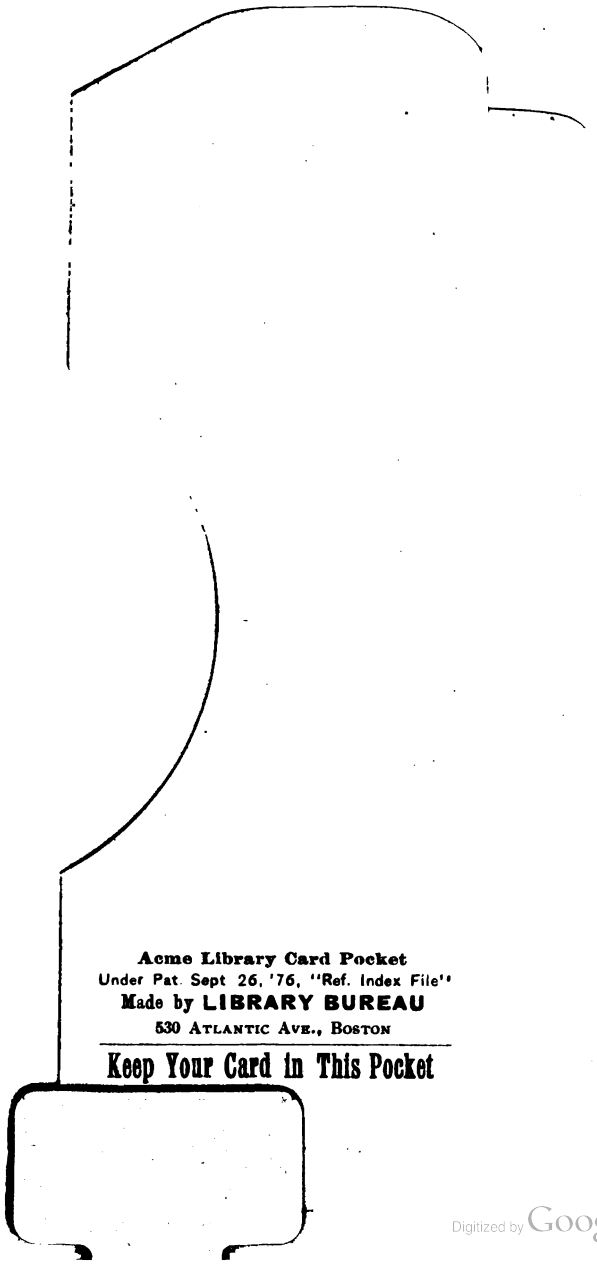

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